

EXPERIMENTS

in Laboratory Chemistry

REVISED EDITION



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L COUKE LOUDEN

MODERN CLASSIFICATION OF THE

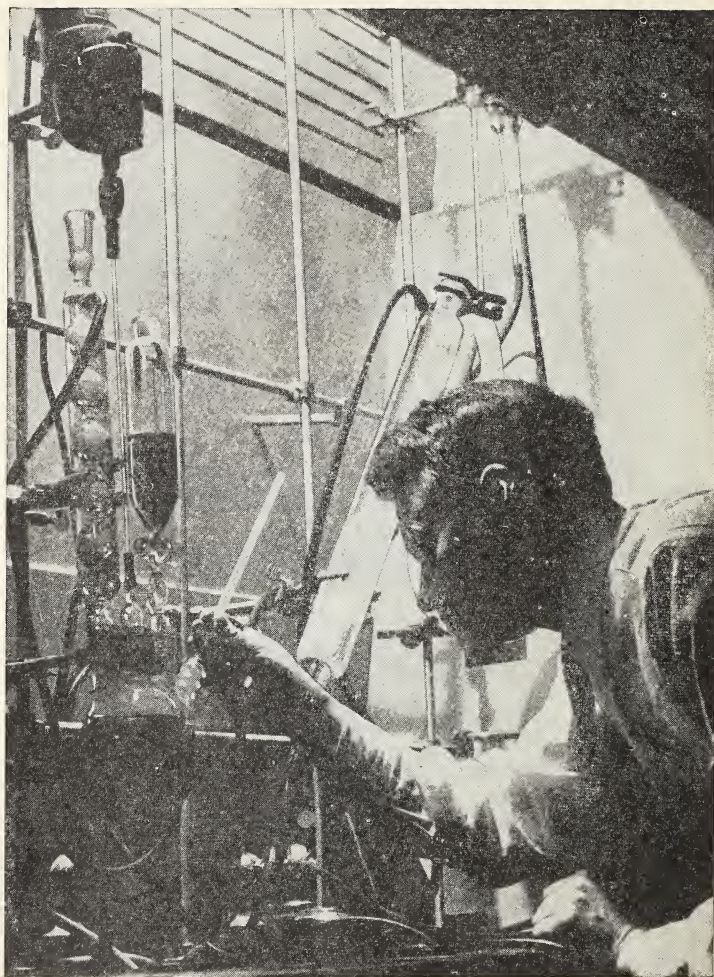
Period	Group				MODERN CLASSIFICATION OF THE															
	I A		II A		TRANSITION															
1																				
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3	2 8 1	11 Na 22.991	2 8 2	12 Mg 24.32																
4	2 8 8 1	19 K 39.100	2 8 8 2	20 Ca 40.08	III B	IV B	V B	VI B	VII B											
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5	2 8 18 8 1	37 Rb 85.48	2 8 18 8 2	38 Sr 87.63	2 8 18 9 2	39 Y 88.92	2 8 18 10 2	40 Zr 91.22	2 8 18 12 1	41 Nb 92.91	2 8 18 13 1	42 Mo 95.95	2 8 18 13 2	43 Tc (99)	2 8 18 15 1	44 Ru 101.1				
6	2 8 18 18 8 1	55 Cs 132.91	2 8 18 18 8 2	56 Ba 137.36	57-71 See Lanthanide Series		2 8 18 32 10 2	72 Hf 178.58	2 8 18 32 11 2	73 Ta 180.95	2 8 18 32 12 2	74 W 183.86	2 8 18 32 13 2	75 Re 186.22	2 8 18 32 14 2	76 Os 190.2				
7	2 8 18 32 18 8 1	87 Fr (223)	2 8 18 32 18 8 2	88 Ra 226.05	89-100 See Actinide Series		NOTE "A" groups consist of elements where the outside ring carries the valence electrons. "B" groups are the transition groups where a shell other than the outermost is being filled. Vertical column to the left of the element shows ring structure. The heavy step line represents the division between metals and non-metals. The heavy vertical line separates the inert elements.													

Lanthanide Series	2 8 18 18 9 2	57 La 138.92	2 8 18 19 9 2	58 Ce 140.13	2 8 18 20 9 2	59 Pr 140.92	2 8 18 22 8 2	60 Nd 144.27	2 8 18 23 8 2	61 Pm (145)	2 8 18 24 8 2	62 Sm 150.35
Actinide Series	2 8 18 32 18 9 2	89 Ac 227	2 8 18 32 19 9 2	90 Th 232.05	2 8 18 32 20 9 2	91 Pa 231	2 8 18 32 21 9 2	92 U 238.07	2 8 18 32 22 9 2	93 Np (237)	2 8 18 32 23 9 2	94 Pu (242)

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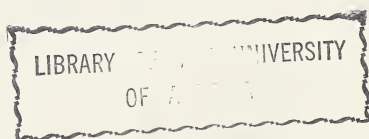


EXPERIMENTS
in
Laboratory Chemistry
Revised Edition



Observe carefully and measure accurately.

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(completely reset)	



Preface

Experiments in Laboratory Chemistry is designed for the beginner in chemistry and is especially intended for use with the text *Chemistry for Secondary Schools* by the same authors. The student is advised to use the text to clarify and co-ordinate the experimental work of the laboratory.

This manual contains more experiments than can be carried out in a year's work. It is suggested that, from the material provided, the teacher will select those experiments that are most suitable for his classes and that can be performed with the apparatus available. It is better to have a smaller number of experiments performed carefully, thoroughly understood, and well written, than to rush through many experiments carelessly. Experiments should not be omitted, however, if their omission in any way interferes with the continuity of the course.

For most of the experiments, only simple inexpensive apparatus is needed. Experiments designated (*Demonstration*) are either too costly, too difficult, or too dangerous to be performed by beginners, and should be carried out on the lecture table, either by the teacher or by assisting pupils under close supervision.

A mere mechanical performance of an experiment has little value. Consequently, the student should become familiar with the background of each experiment by carefully reading the exercise beforehand. Where an experiment is involved or lengthy, the reading assignment should be given on the day previous to the laboratory period.

Complete and precise directions are given which are clearly stated and illustrated, so that the teacher is not burdened with the necessity of dictating details of manipulation and procedure.

Following most procedures is a series of questions (*in italics*) designed to direct the attention of the student to the important features to be observed. Questions also occur at the end of some experiments and units, where it is necessary to gather together the observations of several procedures or experiments in order to make generalizations.

Chemistry is essentially an experimental science. The laboratory work gives reality to the subject by presenting concrete sensory

experiences which aid the student in grasping important ideas, and makes the fundamental principles a part of his own experience. The laboratory is the best place in which to develop a truly scientific attitude.

A. G. C.

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Preface to the Revised Edition

This revised edition of *Experiments in Laboratory Chemistry* has been modified to conform to its companion text, the revised edition of *Chemistry for Secondary Schools*. Additional experiments have been included to satisfy the requirements of most fundamental courses in chemistry.

As stated in the preface to the first edition, the authors are fully aware that more experiments are included than are necessary for the completion of a successful year's work. In this revision, the number of experiments have been increased, so that teachers may make a more varied selection both for class use and for demonstrations.

The authors wish to express their appreciation for the very favorable reception accorded the first edition of *Experiments in Laboratory Chemistry*. We trust that this manual and its companion volume, *Chemistry for Secondary Schools*, will continue to serve Canadian students in building a solid foundation in chemistry.

The authors take this opportunity of expressing their thanks to Mr. R. B. Britton-Foster for his close supervision during the revision and for his many excellent suggestions.

June, 1959.

J. H. C.

A. H. L.

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General Laboratory Instructions

The following suggestions are intended for the guidance of the student:

1. Know your equipment. Familiarize yourself with the name and general purpose of each piece of apparatus in your locker.

2. Use clean equipment. Cleanliness is essential if you are to do accurate work. Experiments performed with dirty equipment often produce misleading results.

3. Be tidy. Before you leave the laboratory, be sure that the gas jets and water taps are turned off, that all your equipment is clean and in its proper place, and that the desk top is sponged and dried. A single drop of acid may cause a severe burn, or damage the clothing of a student in the class that follows. Exempting desks with soapstone tops, all treated or painted desk tops are easily damaged by acids, bases, and many organic solvents. Therefore, a *wet* sponge or towel should always be at hand during experiments. A *dry* sponge or towel should never be used.

4. Use the sink properly. Solids such as burned matches, discarded litmus papers, broken glass, metals, and unconsumed *insoluble* chemicals should never be thrown into the sink, where they may clog the plumbing. Place such articles in the waste jars provided for that purpose. *Soluble* substances and liquids (with the exception of mercury and solutions of mercury compounds) may be poured into the sink, but should be flushed down with plenty of running water. Metallic mercury and solutions of mercury compounds should be poured into a special container provided by the teacher, since they will attack and damage the plumbing.

5. Follow directions. Before beginning an experiment, the directions should be read through carefully until the method of procedure is clearly understood. Then follow the directions closely.

6. Mistakes are dangerous. Always be certain that you have the correct materials for the experiment. Take time to read the labels on reagent bottles carefully. Do not guess the nature of substances contained in unlabelled bottles. Use of the wrong material may cause accidents and painful burns.

7. Quantities of materials. Chemicals are expensive, and the results of experiments are often obscured by using *too much*. You will save time and achieve better results by using the quantities called for in the directions.

8. First Aid. The teacher in charge should be told at once of any accident or personal injury (burn, scratch, cut, corrosive liquid on skin, clothing, desk top, etc.) no matter how trivial it may appear. Also familiarize yourself with the *First Aid Instructions* listed in the appendix (pp. 190-193).

9. Laboratory techniques. In a later section, certain simple operations and manipulations are discussed, which you will be required to carry out in laboratory practice. Possibly the best method is to learn each particular operation at the time it is first encountered in the experimental procedures. This technique should be used thereafter on all occasions.

(*Technique . . .*) has frequently been inserted at various places in this manual to caution both teachers and students to turn to the section on laboratory techniques and familiarize themselves with the particular technique indicated.

10. Recording observations. Concise observations should be recorded as the experiment proceeds.

Laboratory Notebooks

The chief benefits that a student should derive from an elementary course in chemistry are training in careful observation, practice in the proper manipulation of apparatus, and in the accurate description and interpretation of experimental results. The keeping of a good, well organized notebook, therefore, is an essential part of a student's training.

It is not necessary that all experimental records conform to a rigid pattern; rather, after some initial guidance has been given, the student should be encouraged to organize his own notebook. Certainly, the study of chemistry should not be sacrificed to meticulous conformity to mechanical notebook procedure.

Although loose-leaf notebooks that accommodate the various subjects taken by the student are commonly used, many teachers prefer a separate chemistry notebook. A single hard-covered book (approximately 8" $\times$ 11") containing approximately 100 pages of unruled paper is selected by many teachers. This may be fitted with a liner ruled by the student to meet his individual writing style.

The following are a few suggestions to the student:

1. Do not duplicate the printed instructions in the book. Instead, state in your own words, (a) what was actually done, (b) what was actually observed, and (c) the conclusions that were drawn from the facts revealed by the experiment. Be careful to draw only those conclusions that are warranted by the facts.

2. Arrange your records in an orderly manner, and write in ink. Record the number and title of each experiment and the date on which the experiment is performed. Each record should be clear and complete, so that anyone reading it would have no difficulty in understanding the experiment.

3. A well-labelled cross-sectional diagram will make your written description of the experiment easier to understand. The diagrams should be large enough for clarity and understanding. Good judgment should be exercised in the time spent on making diagrams.

4. Answer each *italicized question* with a definite statement, in such a manner that no difficulty will be encountered in recalling the question.

5. Where the experimental work does not provide sufficient information to answer a question adequately, a page reference to the text has been given, indicating where the necessary information may be obtained. The text, however, should not be consulted until the experiment has been performed.

Percentage Error

The difference between your experimental result and the accepted value which has been established by trained scientists making a large number of trials with high-grade equipment is called the *actual error*. A much more significant expression of the degree of accuracy is to find what fraction (expressed as a percentage) this actual error is of the accepted value. For example, if the gram-equivalent weight of magnesium is 12.16 gm. and your experimental result is 11.86 gm., the *actual error* is 0.30 gm., which seems on first inspection to be a sizable error. The *percentage error* is $\frac{0.30 \times 100}{12.16}$, or 2.46%.

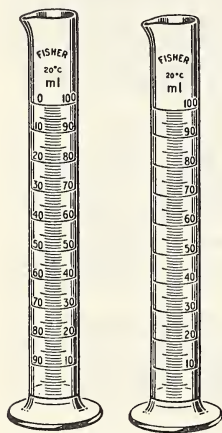
(To find the *percentage error*, divide the *actual error* by the *accepted value*, and multiply the quotient by 100.)

It is impossible to establish a uniform rule for the degree of accuracy to be expected in the various quantitative experiments covered in this manual.

Laboratory Techniques

A. WORKING WITH LIQUIDS

1. Measuring volumes. (a) *Graduated cylinders.* The volumes of liquids are commonly measured in cylindrical glass vessels with graduated scales on their sides. They are often called *chemical graduates* and are obtainable in various capacities. The scale etched in the glass shows the volume in *millilitres*, or in some cases in fluid ounces. Some have a single scale in which the divisions are numbered from the bottom upwards, enabling the observer to measure the volume of liquids *poured into* the graduate. Others have a double scale, the divisions in the second scale being numbered from the top downward, thus making it possible to measure the volume of liquids *poured from* the graduate (Fig. 1).



Fisher Scientific Company

Fig. 1. Graduated cylinders.

You should familiarize yourself with the scale markings on the cylinders available in your laboratory, and their values. You should also note that the top surface of an enclosed liquid is curved, not flat. This curved surface is called the *meniscus*. With liquids such as water, which wet the glass walls, a *concave* meniscus is observed. In reading the volume of such liquids, care must be taken to have the graduate on a fixed, level surface, and the eye must look horizontally along a line passing through the lowest point on the meniscus. Figure 2A illustrates the correct method of reading the volume of a liquid which displays a concave meniscus.

Mercury does not adhere to the glass walls of a graduated cylinder and a *convex* meniscus is observed. In this case, one must be sure that the eye is looking horizontally along a line passing through the highest point on the meniscus.

Mercury does not adhere to the glass walls of a graduated cylinder and a *convex* meniscus is observed. In this case, one must be sure that the eye is looking horizontally along a line passing through the highest point on the meniscus.

(b) *Burettes.* These are calibrated tubes of various capacities, with the zero mark found near the top (Fig. 2B). A glass stopcock near the bottom controls the flow of liquid from the tube. They may be filled by placing a funnel in the top of the tube and adding the liquid until it rises above the zero mark. The stopcock (tap) is then opened and sufficient liquid is allowed to escape until all the air in the tip has been driven out and the upper level of the liquid is *at*

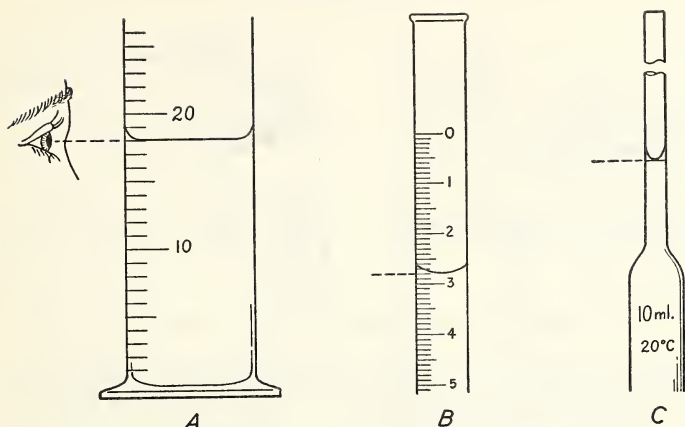


Fig. 2. Measuring volumes of liquids. A, using a graduated cylinder; B, using a burette; C, using a pipette.

or *below* the zero mark. Any desired volume may then be released by careful manipulation of the tap. (TEXT Fig. 16.4) The initial and final positions of the meniscus are read in the manner described earlier.

(c) *Pipettes*. The pipette is a glass tube tapered to a point at one end and commonly having a bulb-like enlargement near its centre. It is calibrated to deliver a specified volume of liquid at the temperature indicated. The pipette is filled by suction until the liquid is slightly higher than the graduation mark found on the stem. Then the upper end of the tube is quickly covered with the finger. By a slight release of the pressure exerted by the finger, the liquid is allowed to descend slowly until the bottom of the meniscus curve is coincident with the mark on the stem, as shown in Figure 2C. In allowing the pipette to deliver its specified volume of liquid, *allow it to drain freely*. Do not blow through it or aid its draining in any way.

CAUTION: Pipettes should not be used to dispense poisonous or volatile liquids. If in doubt, a suction bottle should be used to fill the pipette.

2. Using reagent bottles. When pouring a liquid reagent from a stoppered bottle, *the bottle or its stopper must never be placed on an unprotected desk surface*. Since most desk surfaces are not immune to damage, many protective devices (glass plates, glass or enamelled saucers, transite squares, etc.) are used at the discretion of the teacher in charge, and the student should closely observe the routine established in his particular school.

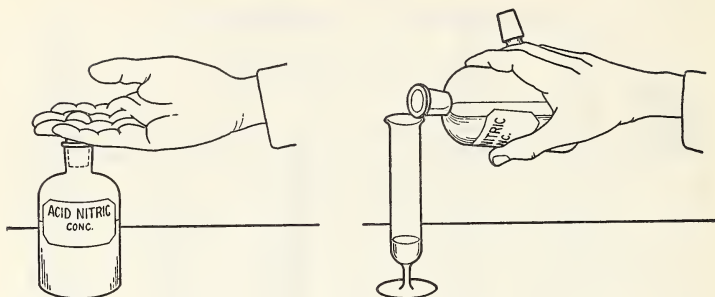


Fig. 3. Pouring a liquid reagent from its container.

When pouring a liquid reagent, the stopper should be held as shown in Figure 3 to avoid contamination. If a contaminated stopper is returned to the bottle, it may make the contents unfit for further use. For the same reason, students should never pour excess reagents back into bottles or exchange the stoppers of bottles.

When pouring, the neck of the bottle should *touch* the rim of the receiving vessel. This prevents the liquid from trickling down the side of the bottle and is especially important when handling acids, bases, or other corrosive liquids. There are many reagents a single drop of which, carelessly left on the desk top, can mar the finish, damage someone's clothing, or cause painful burns.

Bottles containing acids or volatile organic liquids should never be stored near radiators, or allowed to stand in direct sunlight, since dangerous gas pressures may be built up.

3. Diluting concentrated sulphuric acid. In preparing a dilute solution of sulphuric acid **never pour water into the concentrated acid.** *The concentrated acid should be slowly added to the water in a thin stream accompanied by constant stirring.* When concentrated sulphuric acid and water are mixed, a great quantity of heat is evolved. If the water is poured into the acid, the heat generated may convert some of the water into steam, and this may cause some of the hot acid solution to be violently ejected.

4. Filtering. Insoluble material may be separated from a liquid by filtering. The porous filter paper is prepared for the funnel by folding it successively into the shapes shown in Figure 4. When the paper has been folded along its diameter and then again at a right angle to the first fold, one of the outside folds is opened so that the paper assumes a cone-like shape with three thicknesses of paper on one side and one thickness on the other. This is placed in the funnel and sufficient

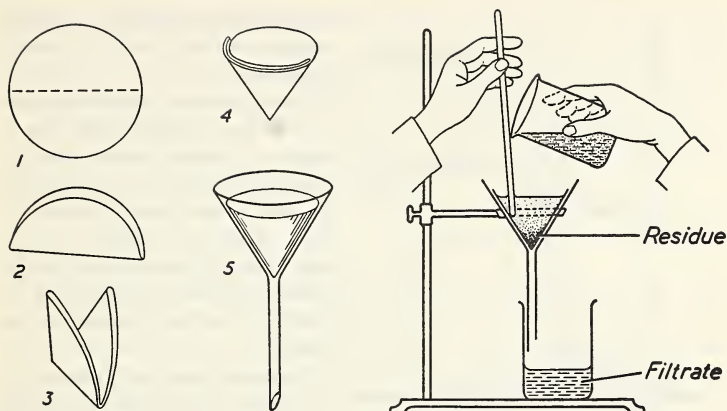


Fig. 4. The process of filtration.

water added to make it adhere to the glass (with no air spaces), so that it will filter rapidly. This treatment also discourages the absorption of large quantities of the solution being filtered. After the moist paper has been pressed down with the fingers, any excess water should be discarded before using the filter. The paper should not extend above the edges of the funnel, but its apex should always project slightly into the stem.

Support the funnel in a ring attached to a retort stand and place an empty beaker or test-tube under the funnel to catch the *filtrate*. The liquid to be filtered should be poured slowly down a glass rod into the funnel, as shown in Figure 4. The lower end of the rod should touch the filter paper inside the funnel, so that the liquid will run down the side and thereby avoid breaking the apex of the filter paper. The acute edge of the funnel stem should touch the side of the collecting vessel. This permits the filtrate to run down the side of the receiving vessel and prevents spattering.

After the liquid has drained from the solid material (*residue*) retained by the paper, soluble impurities are often removed by *washing* the residue with several 5 ml. portions of distilled water, and allowing each portion to run through before adding the next (**Technique 9**).

5. Evaporating over a water bath. If the evaporating dish containing the solution to be evaporated is placed on a wire gauze supported on a retort stand, and slowly heated over a low flame, "spattering" commonly occurs as the condition of dryness is approached. This causes the loss of some of the final product and may be minimized by

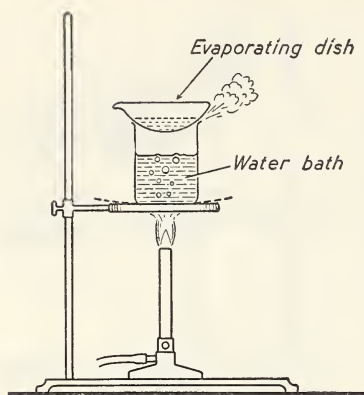


Fig. 5. Using a water bath.

covering the dish with a glass plate or watch-glass of appropriate size.

Usually it is advisable to support the evaporating dish over a beaker of boiling water as illustrated in Figure 5. The dish is heated by the steam produced. Care must be taken to replace the water in the beaker from time to time, since allowing it to boil dry is likely to crack the glass.

In some non-quantitative experiments where only a small volume of a solution is to be

evaporated to dryness, the "water bath" method is not used. Instead, the evaporating dish containing the solution is supported on wire gauze and is heated strongly at first until the volume of liquid is considerably reduced. Then, the final evaporation to dryness is accomplished with a much lower flame, since small pockets of steam cause the liquid to "spatter" as the solution becomes more concentrated.

When only a few drops of a solution are to be evaporated, a watch-glass is usually employed. This should be encircled with the thumb and first finger and held over a low flame. Unless it is held thus, it may be brought too close to the flame and the watch-glass may crack.

6. Testing with litmus. Except where a different method is specifically stated, litmus tests on *liquids* should be conducted by placing the paper (or papers) on a clean glass plate and transferring a drop or two of the liquid to the paper with a clean glass rod.

When *gases* are to be tested, the litmus paper (or papers) should first be moistened with distilled water.

Never handle litmus papers with wet or dirty fingers, nor permit them to lie on the desk top.

B. WORKING WITH SOLIDS

7. Heating glassware. (a) Every piece of apparatus made of glass or porcelain should be *thoroughly dry on the outside* before being heated.

(b) *The temperature should be raised gradually.* Sudden heating or cooling, or concentration of the flame at one spot for any length of time, may cause the apparatus to crack.

(c) Beakers, flasks, evaporating dishes, glass retorts, etc., should not

be heated in direct contact with the flame. Support them on a wire gauze with an asbestos centre. A porcelain crucible and its lid, which have been warmed gradually in the initial heating and allowed to cool slowly, may be supported on a clay triangle and heated directly.

(d) Never clamp flasks or test-tubes more firmly than is necessary to hold them in position. The expansion of the glass on being heated may break it if it is clamped too tightly.

(e) Test-tubes may be heated with a direct flame. A strip of paper several times folded, may be substituted for a wire holder if these are not available (see Fig. 23, p. 36).

In heating a test-tube containing a liquid, it should always be held in an *inclined position*. The temperature should be raised *gradually* and *uniformly* to avoid the sudden ejection of the contents because of chemical action or the formation of steam. This uniform heating may be accomplished by moving the tube back and forth through the flame, heating it from the upper portion downwards, and gently agitating the contents. **Always direct the mouth of the test-tube away from yourself or anyone near you, while the heating is in progress.**

A note to teachers: The authors believe that the use of *soft-glass* test-tubes in a high school laboratory is a dangerous and undesirable policy. *Pyrex* test-tubes are more expensive but give much longer service and will cause fewer breakages and accidents. The authors also believe that Pyrex flasks (Florence or Erlenmeyer types) should always be used as generators instead of glass bottles, thereby avoiding breakages resulting from the exothermic character of many reactions.

8. Putting a solid into a test-tube. Cut a piece of stiff, glazed paper an inch longer and slightly wider than the diameter of the test-tube used. Fold it lengthwise to form a V-shaped trough. Place the solid near one end of the trough. Hold the test-tube in a horizontal position and slide the trough inside the tube, as shown in Figure 6. By inclining the tube and gently tapping the paper, the

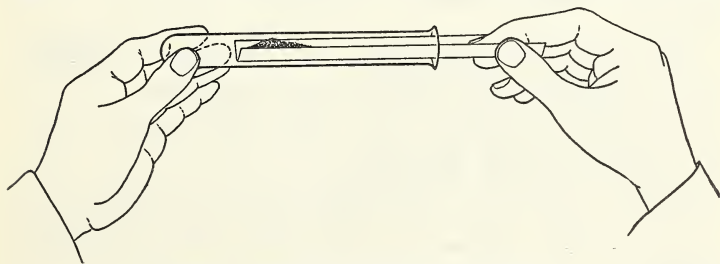


Fig. 6. Introducing a solid into a test-tube.

material will be deposited in the bottom of the test-tube, leaving the sides clean.

9. Washing a precipitate. When an insoluble substance is filtered out of a solution, the pores of the filter paper retain a portion of the solution itself. Some of the solution also adheres to the particles of retained solid. To eliminate all traces of the soluble material, the residue is washed by adding distilled water. The wash bottle shown in Figure 7 is commonly used as a container for the distilled water. On blowing into the mouth-piece, a fine stream is forced out of the jet. Sometimes the water is poured directly from the mouth-piece in 5 to 10 ml. portions, and each portion is allowed to run through the filter paper before adding the next. This treatment is continued until all the soluble

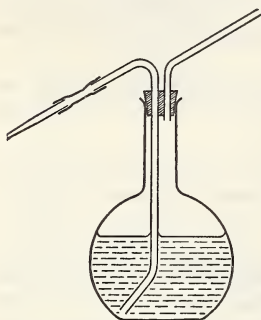


Fig. 7. A wash bottle

material is washed through. The insoluble material can then be dried and weighed.

10. Handling white (yellow) phosphorus. White phosphorus is stored under water because it has a low ignition temperature and takes fire on exposure to air. It causes severe burns if allowed to come in contact with the skin. Consequently, white (yellow) phosphorus must always be cut under water and handled with forceps; never with the fingers.

If cut fragments are made available at the teacher's table, the student must keep the phosphorus under water while carrying it to his desk.

Any traces of unburned phosphorus remaining on a deflagrating spoon (Fig. 8.) must be burned before putting the spoon away. This is done by heating the spoon to redness for two minutes. Phosphorus deposits on flasks and test-tubes may also be very safely removed by being immersed in a solution of cupric sulphate or in an acidified solution of bleaching powder.

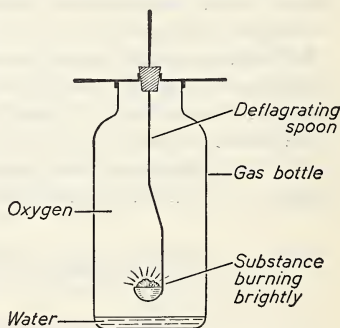


Fig. 8. Using a deflagrating spoon.

11. Inserting glass tubing into stoppers or rubber tubing. It must be remembered that glass is fragile and that carelessness in this simple operation can easily result in a badly lacerated hand. Always *lubricate* the end of the glass tubing with soap solution, vaseline, or glycerine before introducing it into the stopper or rubber tubing. Hold the stopper or rubber tubing between the thumb and the forefinger of one hand, and with the other hand, grasp the glass tubing *with a towel* at a point approximately one inch from the entering end. Always aim the glass tubing *away from the palm of the hand* which holds the stopper or rubber tubing. Use a *gradual rotary motion* with the towel-wrapped hand and shift the grip as often as necessary.

C. WORKING WITH GASES

12. Collecting gases by displacement of water. When gases are to be collected over water, the gas bottle should be filled with water until the rounded upper surface of the water extends above the mouth of the bottle, as shown in Figure 9A. A glass plate should then be slid over the mouth of the bottle and held tightly in place while the bottle is being inverted over water in a pneumatic trough or the sink. The glass plate is removed as soon as the mouth of the bottle is below the surface of the water in the trough.

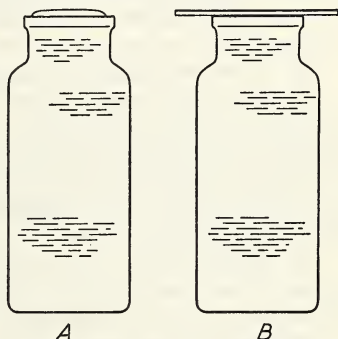


Fig. 9. The proper method of filling a gas bottle with water in collecting a gas by the downward displacement of water.

To collect the gas, the bottle should be tipped enough to permit placing the end of the delivery tube under its mouth. When rubber tubing is used as the delivery tube, care must be taken not to pinch the tubing with the tipped bottle. The use of a "beehive" support, as shown in Figure 10A, eliminates the need for tipping the bottle, since it has an opening in the top over which the water-filled gas jar is rested. An arched opening in the base permits the insertion of the tubing through which the gas passes. The pneumatic trough provided in your laboratory may be equipped with a removable, perforated, sliding shelf, which functions in the same manner as a beehive support.

In experiments where the flask or test-tube in which the gas is being generated is heated, care must be taken never to allow the

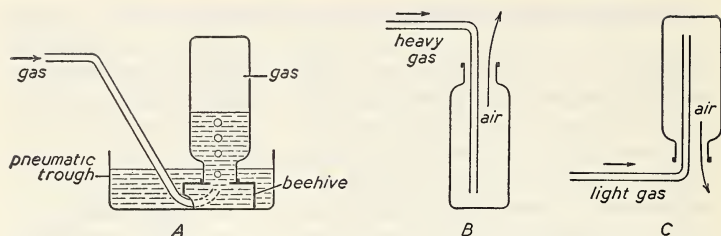


Fig. 10. Three methods of collecting a gas. *A*, downward displacement of water. *B*, upward displacement of air. *C*, downward displacement of air.

generator to cool while the free end of the delivery tube is still under water. As the generator cools, an internal decrease in pressure occurs. The cold water forced back into the generator by the greater atmospheric pressure may crack the generator.

13. Collecting gases by displacement of air. If the gas being collected is appreciably soluble in water, collection by displacement of water has to be discarded. If the gas is heavier than air, it is collected by the *upward* displacement of air, as shown in Figure 10*B*. If the gas is lighter than air, it is collected by the *downward* displacement of air, (Fig. 10*C*). The delivery tube should be extended to the base of the collecting bottle in both cases.

14. Smelling a gas. Use care in smelling gases. Keeping your mouth closed, wave the gas toward your nose with your hand, and sniff gently. **Never inhale gases.**

15. Making splint tests. Wooden splints of cedar or pine are frequently used to determine whether a particular gas either burns or supports combustion. They should be 8 to 10 inches long, and about $\frac{1}{4}$ inch in diameter. Some experiments demand the use of a *glowing* splint; in others a *blazing* or *flaming* splint is required. A glowing splint is prepared by igniting the splint in a Bunsen flame and allowing it to blaze until there is a spark of glowing charcoal at the tip when the flame is shaken out.

Burners for Gaseous Fuels

The Bunsen-type burner is almost indispensable as a source of heat and is used in all laboratories where some type of gaseous fuel is available. Various models are available, but all are so constructed that a mixture of air and the gaseous fuel can be ignited at the top of the vertical barrel. The nature and temperature of the flame are controlled by the relative volumes of gas and air admitted.

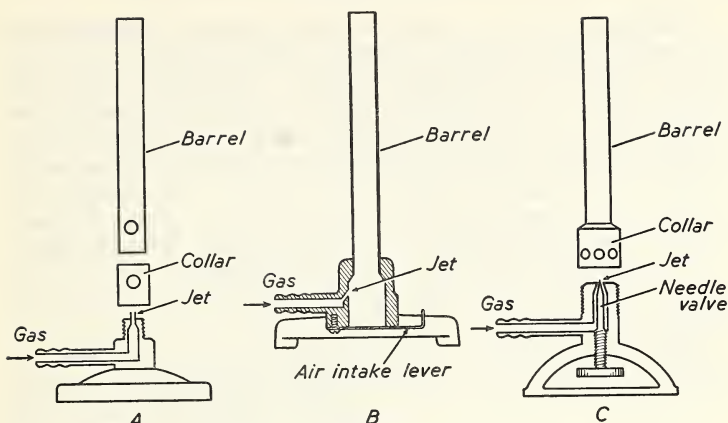


Fig. 11. Various types of burners.

The burner represented in Figure 11A is a typical Bunsen burner. The barrel may be unscrewed from the base, thus revealing the small opening in the jet through which the gas enters the mixing chamber. The collar is a perforated ring which fits over the base of the barrel in such a way that the openings in the barrel may be kept open or closed by rotating the collar. The burner is connected to a gas outlet by means of tightly fitted thick-walled rubber tubing.

Model B has an H-shaped base, and the orifice which admits the gas is displaced to the side wall of the stack, the latter being firmly seated in the base casting. An air-intake adjustment of special design is operated by a lever on the underside of the base.

In model C, the air inlets are located in the base of the barrel and the gas flow is controlled by a knurled adjustment screw which regulates a needle valve.

In some models, the gas chamber, orifice, air jet, and hose connection are all contained in the base casting and a slight rotation of the barrel adjusts the air supply. Possibly your laboratory may have *Meker* or *Fisher* type burners in which the top of the barrel is fitted with a metallic grid. The latter prevents back-flashing and produces a flame that is quite different from the Bunsen flame. Instead of the long narrow Bunsen flame with a hot tip and a cooler central cone, a wide flame of great intensity is produced, which is uniformly hot throughout.

Examine the burner provided in your laboratory. Note the features of its construction, and make a labelled diagram of the burner,

Lighting the burner. Attach the rubber tubing of the burner to a gas outlet, and control the air supply by *closing* the air inlet at the base. Light the match before opening the gas jet, and after a few seconds, light the burner by bringing the match flame across the top of the barrel. Do not hold the lighted match over the burner before turning on the gas, since this procedure will probably cause the match flame to be extinguished before the gas ignites. Then gradually increase the air supply and adjust the gas supply to obtain the high or low flame required for the particular experiment to be performed.

If wax tapers or gas lighters equipped with spark metal tips are used to eliminate the use of matches in the laboratory, the same instructions apply.

High and low flames. Since the burner operates most efficiently when the mixture of air and gas is such that the flame is neither

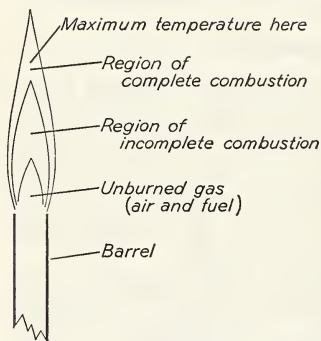


Fig. 12. The various regions in a Bunsen flame.

coloured yellow nor noisy, you should adjust the air supply until all traces of yellow disappear, but not enough to make the flame noisy. Establish a quiet, non-luminous (blue) flame displaying the distinct cones shown in Figure 12. This is the type of flame you will employ in your experiments unless directed otherwise.

The region closest to the burner barrel represents an unburned mixture of air and the fuel gas. If an object is to be heated to best advantage, it should be held *just*

above the tip of the central cone, since the maximum temperature of the Bunsen flame occurs at this point. If a *high* flame is specified, the tap on the gas supply should be regulated until a flame not more than 4 or 5 inches high is obtained. When a *low* flame is specified, the gas supply should be reduced.

Learn to create a high flame displaying the zones shown in Figure 12.

Draw the upper portion of the barrel of the burner used in your laboratory, and show a properly adjusted high flame with its cones.

Label the various zones and indicate where an object should be placed if it is to be heated to best advantage.

Also learn how to make a low flame.

A luminous flame. When the air supply is inadequate, the flame produced is sooty, luminous, and not very hot. Cut off the air supply

on your burner by closing the air inlet, and note the yellow flame produced. Using a pair of tongs, hold a porcelain crucible cover or a glass plate in this flame. The black deposit is carbon, resulting from the incomplete burning of the gas, and the yellow colour of the flame is due to the heating to incandescence of unburned carbon particles. You should never use this type of flame unless specifically directed to do so.

Back-flashing. If too much air is admitted, the flame sometimes "*strikes back*" and burns at the jet near the base of the burner. When this happens, turn off the gas supply at once and let the burner cool. Then *decrease* the amount of air admitted to the burner and relight. Back-flashing can also be controlled by adjusting the flow of gas entering the burner. Back-flashing is dangerous because of the poisonous gases released in the laboratory. It can also cause severe finger burns if the burner is touched.

Working with Glass

Cutting. Determine the length needed, lay the tubing on the desk, and at the place where it is to be cut make a short but deep transverse scratch with a forward stroke of a triangular file. Then grasp the tubing firmly with a hand on either side of the scratch, holding the thumb tips together behind the mark. Now press outward with the thumbs, pulling gently inward at the same time with the hands, and the tube should break at the point where the scratch was made (Fig. 13). If the tubing does not yield readily to gentle pressure, a deeper scratch must be made.

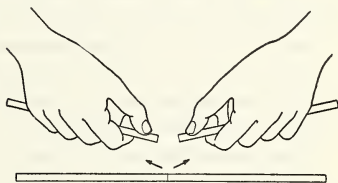


Fig. 13. Cutting and breaking glass tubing.

The sharp edges should be rounded off (fire-polished) by slow rotation at a downward angle of 45° in a Bunsen flame until the glass begins to melt slightly. Relatively smooth edges may be rubbed with emery cloth or a file instead of being fire-polished.

Bending. Glass tubing should always be bent in a *flattened* flame. A Bunsen flame can be flattened by the use of a "wing top" attachment which slips over the top of the burner tube, as shown in Figure 14. For ordinary bends the flattened flame should be slightly yellow and about two inches in width.

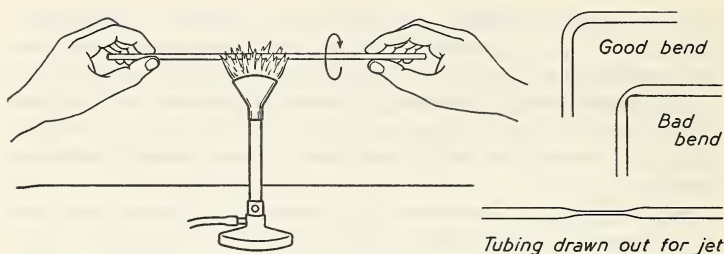


Fig. 14. Heating glass tubing.

A *right-angle* bend is made as follows. Determine the point at which the tubing is to be bent. Grasp the tubing in both hands and hold it slightly above the flattened flame so that about 5 or 6 centimetres are heated where the bend is to be located. Slowly *rotate* the tubing between the thumbs and forefingers to insure uniform heating and gradually lower it into the position shown in Figure 14. Continue the slow rotation until the tube is hot enough to sag of its own weight. Then *remove it from the flame* and immediately bend it to the desired angle. A block of wood or a similar right-angled object resting on an asbestos mat is of great value in achieving an exact right angle and in ensuring that all portions of the tubing will lie in the same plane. In cooling, the hot glass must not be allowed to come in contact with the surface of the desk or any cold object.

If angles other than right-angles are desired, the bending is simplified if the desired angle is previously drawn in pencil on the asbestos mat. The heated tubing is then made to coincide with this angle.

Making a glass jet. Discard the "wing top" and heat the tubing in a non-luminous flame. Rotate the tube to insure uniform heating. When the heated region is soft and begins to sag, *remove from the flame* and slowly draw the ends apart in a straight line until the constricted portion is of the desired diameter. Allow the tubing to cool, make a file scratch at the point where it is to be severed, and break. Then fire-polish the cut end, being careful not to seal it.

Inserting glass tubing into stoppers and rubber tubing. *Carelessness in this simple operation can easily result in a badly lacerated hand.* Be sure that the ends of the tubing are properly fire-polished. Lubricate both the tubing and the opening in the stopper with soap solution, vaseline, or glycerine. Hold the stopper between the thumb and forefinger of one hand and grasp the *towel-wrapped* tube with the other hand about one inch from the end closest to the stopper. Slowly work the tubing into the opening by a gradual rotary motion.

Never grasp any bent tubing at the bend when inserting it into a stopper, and never point tubing toward the palm of the hand which holds the stopper.

Constructing a wash bottle. A wash bottle is shown in Figure 7. The nozzle should be drawn out so that only a moderately fine stream of water is forced out by gently blowing into the mouthpiece. All the ends of the glass tubing used must be fire-polished. Wash bottles are of value as containers for distilled water to be used in washing precipitates, making special solutions, rinsing glassware, etc.

Working with Balances

A. A CHEMICAL BALANCE

The chemical balance is a delicate instrument that will weigh light objects to a high degree of accuracy and is of inestimable value to the chemist. Since the quantities of substances to be weighed are usually small, these balances have to be much more accurate than those in ordinary use in the commercial world. If one recalls that one ounce is equivalent to 28.35 grams and that in quantitative experiments one has to make accurate weighings in terms of tenths, hundredths, or even thousandths of a gram, it is apparent that this sensitive instrument must be used with a great deal of care, since it may be easily damaged by careless manipulation.

Figure 15 represents an inexpensive simple model. A vertical pillar (P) is firmly attached to a wooden base. The metal beam (B B_1) acts as a lever and has a central wedge-shaped knife-edge which points downward and rests upon a true surface or bearing-plane situated at the top of the pillar, when the balance is in the weighing position. This permits the beam to swing freely on the fulcrum with a minimum of friction. Some models have a V-shaped groove at the top of the pillar instead of the flattened bearing-plane. A knife-edge is also found at each end of the beam (E_1 and E_2). These are turned upwards and support the stirrups (S and S_1) by fitting into V-shaped grooves with which the stirrups are provided. All the knife-edges and bearing-planes are constructed from polished agate or hardened steel. Each stirrup has a notched hook which holds a pan support (H and H_1), which carry the removable scale-pans (W and W_1).

When the balance is not in use, the beam is supported by two fixed arresting screws (A and A_1), thus relieving the central knife-edge of needless pressure and possible damage. With the beam

supported by this arresting mechanism, the scale-pans rest on the wooden base and no stress is imposed on their particular knife-edges. The beam is raised and lowered by means of a knurled knob (K) located in the front of the balance. When turned to the right, the beam is lifted from the arresting screws and begins to swing, since it is now in the weighing position, resting on the central knife-edge. In this position, the stirrups are also suspended from their respective knife-edges. When the knob is turned in the opposite direction, the beam is lowered and comes to rest on the fixed screws which are intended to support it when not in use.

Attached to the beam is a pointer (I), which is free to move over a scale (X), thus indicating when a balance is attained between the object to be weighed and the weights used. If the two arms of the balance are in exact equilibrium, the pointer will swing equally on both sides of the middle of the scale. If the swing is incorrect it may be adjusted by turning one of the small balancing nuts (N and N_1) in the required direction. A faulty swing may also be due to the fact

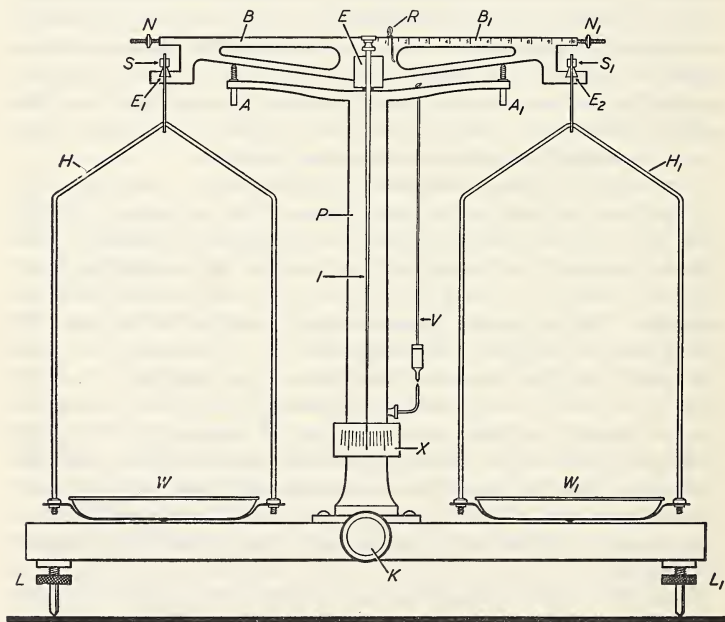


Fig. 15. A chemical balance

that the balance is not level. Two levelling-screws (L and L_1) on the base may be turned if a spirit level shows that the base is not level. A plumb-line (V) at the side of the pillar is sometimes used as a levelling device.

The sliding weight or rider (R) is a looped wire of known weight, bent so that it can be placed astride the beam. It may be moved along the beam when weights of less than 0.01 gm. are required. In the balance (Fig. 15), the rider may be moved along the right-hand side of the beam. In more expensive models, graduations occur on both sides of the centre, so that the weight of the rider can be subtracted as well as added.

A good balance should be housed in a cabinet with clear glass windows on the sides and top and fitted with a counterpoised sliding front door. This glass case is designed to exclude dust and laboratory fumes, and to avoid disturbing air currents.

The student is advised to sit in front of a chemical balance, read the description a second time, and identify the different parts by reference to Figure 15.

Sets of metric weights. Examine the metric weights available in your laboratory and observe how they are marked and arranged. You will probably find:

(a) Brass weights (**grams**): 100 gm., 50 gm., 20 gm., 20 gm., 10 gm., 5 gm., 2 gm., 2 gm., 1 gm.

(b) Aluminum or nickel fractional weights (**milligrams**):

500 mg. (0.5 gm.)	20 mg. (0.02 gm.)
200 mg. (0.2 gm.)	10 mg. (0.01 gm.)
200 mg. (0.2 gm.)	5 mg. (0.005 gm.)
100 mg. (0.1 gm.)	2 mg. (0.002 gm.)
50 mg. (0.05 gm.)	2 mg. (0.002 gm.)
20 mg. (0.02 gm.)	1 mg. (0.001 gm.)

The box also contains a pair of forceps which must be used when weights are removed from, or replaced in, the box.

You should check to ensure that all the weights are present and report any missing weights.

Make a list (or a diagram) showing the weights provided in your laboratory.

B. A TRIAL WEIGHING

Possibly your instructor will want you to do a practice weighing using some object such as a penny, a nickel, a dime, or a small beaker, etc. If such be the case, proceed as follows:

Check the swing. See that the stirrups are not displaced and that the scale-pans are clean and dry. Turn the operating knob, thus allowing the beam to swing, and observe the excursions made by the pointer. If the balance is in proper adjustment, the pointer will travel over an equal number of divisions on each side of the central graduation on the scale. If the pointer travels farther in one direction than in the other, return the beam to its position of rest by turning the operating knob, and adjust one of the small balancing nuts in the proper direction. Again observe the excursions of the pointer when the beam is raised. Repeat the adjustments in this manner until the swings to right and left are equal. It is not necessary to wait for the pointer to come to rest. A good balance will continue to swing for a long time.

Use the metric weights. *Never add or remove anything from the scale-pan while the beam is raised.*



Fisher Scientific Company

Fig. 16. A box of metric weights.

Place the object on the *left*-hand pan. Select the largest weight in the box and *using forceps* place it on the *right*-hand pan. Now partially raise the beam by turning the operating knob. If the weight selected is too heavy, the scale-pan holding the object will rise and the pointer will swing to the left; a movement of the pointer to the right indicates that the weight is too light. Let us assume that the first weight selected is too heavy. Return the beam to its position of rest, remove the weight and return it to the box. Try the next largest weight, and continue in this manner from the largest to the smallest, without omitting any, until a weight is added which

is lighter than the object. Leave this on the pan, add the next lighter weight, and so on through the brass weights. Then work through the fractional weights (milligrams) in the same systematic diminishing order of magnitude, until the pointer swings equally (or nearly so) in both directions.

If your balance is equipped with a rider, the latter is commonly used when weights of less than 0.01 gm. are required. This feature varies so widely in different models that the manipulation of the rider on your particular balance will be explained by the teacher.

The balance case should be closed when observing the movements of the pointer in the final stages of the weighing.

Record the weights. When you have completed the weighing, write down the separate weights carefully before they are removed from the scale-pan and total them at once. This record should be made in your notebook and never upon a loose sheet of paper, since the loss of the latter will necessitate a repetition of the weighing, and often the repetition of the entire experiment. Then, beginning with the largest weight, return them to the box and check the individual weights as you return them. In this way any error will be observed and corrected.

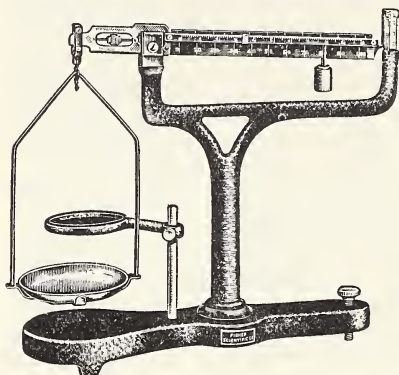
Special instructions. Most of the rules to be observed in the care and manipulation of the balance have been stressed at appropriate places in the earlier paragraphs. The following additional precautions must be observed:

1. Dust must be removed with a "camel's hair" brush.
2. The object to be weighed must be clean and dry.
3. Do not weigh an object when hot; the heat causes warm air currents which affect the weighing.
4. Chemicals are never placed on the unprotected scale-pan. Solids are placed upon a watch-glass or a piece of paper. Liquids are usually placed in a small beaker. The empty container is weighed, and the weight of the substance is determined by subtraction.
5. Set the object to be weighed in the *centre* of the left-hand scale-pan. The larger weights should also be placed in a central position in their pan.
6. When a weighing is completed, be certain that the beam is left in the resting position and that the lid of the weight box is closed.

C. A TRIPLE BEAM BALANCE

This type of balance has three separate graduated beams, mounted in parallel in the same horizontal plane and joined to an arm bearing the hardened steel knife-edge, stirrup, and adjusting thumb-screw.

The front beam is graduated to 1 gram in tenths and hundredths of a gram divisions; the centre beam is graduated to 100 grams in 10 gram divisions; the rear beam is graduated to 10 grams in 1 gram divisions. The front beam has a sliding rider weight, while the other two are notched and are equipped with non-removable weights. A beam release mechanism controls the freedom of movement of the beam.



Fisher Scientific Company

Fig. 17. A triple beam balance.

The base has an adjustable screw for levelling the balance, and many models are equipped with a spirit level as well. A metal rod, attached to the base, acts

as a holder for an adjustable beaker support for specific gravity determinations.

In some types, the free end of the beam is constructed to hold a weight hanger which is used when the load being weighed exceeds 111 grams. The weight of the hanger is 100 grams, and slotted auxiliary weights are provided to fit this hanger, thus increasing the weighing capacity.

D. A TRIP BALANCE

This is a heavy-duty balance used when sensitivity is not of primary importance, as in dispensing chemicals, weighing large samples for making solutions, determining specific gravities, etc.

The two pans are non-corrosive vitrolite plates. The knife-edges are usually hardened steel with self-centering agate bearings. They usually have a single beam, graduated 0 to 10 grams in 0.1 gram divisions. A permanently attached rider eliminates the need for 10 gram and smaller weights. Some

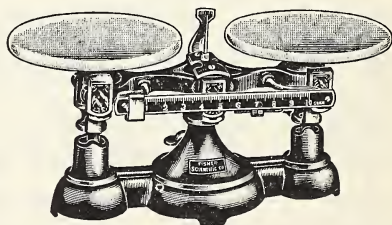


Fig. 18. A trip balance.

models are equipped with a second beam graduated 0 to 200 grams in 10 gram divisions, thereby extending the range over which weighings may be made without accessory loose weights.

In using this type of balance, the student should be certain that the sliding rider is at the extreme left, or at zero. A light touch on one of the pans will set the balance swinging and the pointer should swing at least two or three divisions to either side of centre on the scale. If it swings farther to one side than to the other, a slight turning of the adjusting thumbscrew should be made until the pointer moves through as many divisions to the left as to the right. It is unnecessary to wait until the pointer comes to rest.

The object to be weighed must be placed on the *left* pan. Since jarring dulls the knife-edges and affects the sensitivity of the balance, it should be avoided by supporting one pan with the left hand when weights are being added or removed. The weight of the object is equal to the sum of all weights present on the right pan, plus the weight represented by any movement of the sliding rider from its zero position.

UNIT 1

A Study of Matter

EXPERIMENT 1. The measurement of volume.

1. Read the numbers on a graduated cylinder. Most graduates are marked in millilitres. One litre is defined as the volume occupied by one kilogram of air-free (boiled), distilled water at 4°C. This volume is 1000·028 cubic centimetres, so that 1 ml. is equal to 1·000028 cc. For practical purposes the millilitre and the cubic centimetre may be considered to be the same. Because most vessels used in the laboratory to measure the volumes of liquids or gases are graduated in millilitres, the millilitre will be used as the unit of volume in the experiments which follow.

What is the capacity of the graduated cylinder being examined?

What does each division represent?

2. In measuring the volume of a liquid which wets the walls of the graduated cylinder and forms a *concave* meniscus (*Technique 1a*), place the cylinder on a fixed, level surface and look horizontally along the *bottom* of the meniscus curve (Fig. 2A).

Determine the capacities of several test-tubes of different sizes by filling them with water and measuring the volumes in a graduated cylinder.

From the graduated cylinder, pour 10 ml. of water into an empty test-tube. Stick a narrow paper label on the tube in a position that will permit you to put a pencil mark opposite the water meniscus. Reserve this test-tube for use in later experiments in which you are asked to approximate certain volumes. For example, if you are asked to use approximately 5 ml. of a liquid reagent, it is not necessary to measure this volume precisely. Simply add sufficient liquid to bring the level half way to the pencil mark. With practice you will learn to estimate different volumes fairly accurately. In estimating quantities of solids, the teacher may use spoons of designated sizes to measure out desired quantities rather than to weigh them as suggested in the instructions.

Why is the eye always placed on a line horizontal to the bottom of the curved surface of the liquid whose volume is being measured? Why is the graduated cylinder always placed on a table?

Why do you read from the bottom of the meniscus?

Record the volumes of the test-tubes you have used.

3. Examine a burette (*Technique 1b*). Support the burette on a retort stand by means of a burette clamp, and partly fill it with water. Take a reading by observing the bottom of the meniscus. Record the reading.

Does the reading you have taken represent the volume of water in the burette? Explain.

Do the graduations read from the bottom to the top, or from the top to the bottom?

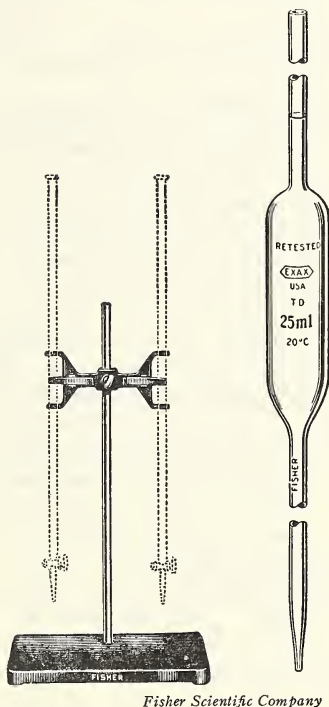
Place a beaker below the burette and allow some water to run into it. Read the burette and record the reading.

What volume of water is in the beaker?

4. Examine a pipette (*Technique 1c*). Place the pointed end of the pipette in water and the opposite end in the mouth, and by suction, draw water into the pipette well above the mark on the stem. Then quickly place the index finger over the upper end of the pipette and at the same time withdraw the pipette from the mouth. If the water level drops below the mark on the pipette the operation must be repeated. Gently ease the finger and allow the water to drip slowly until the bottom of the meniscus is even with the mark on the pipette. Now allow the water to flow freely into a beaker.

What is the volume of the pipette used? Find out the volumes of the pipettes that are available in your laboratory.

Why is there danger when a pipette is used to measure the volume of a poisonous liquid? How could you avoid this danger and still use a pipette?



Fisher Scientific Company

Fig. 19. Burette clamp supported on a stand and holding two burettes. On the right, a 25 millilitre pipette.

QUESTIONS

1. Compare the uses of a graduated cylinder, a burette, and a pipette under the headings of (a) convenience, and (b) accuracy.

2. Make diagrams showing (a) a graduated cylinder containing 47.5 ml. of water, (b) a graduated cylinder containing 38 ml. of mercury, (c) a burette with the meniscus at the 60 ml. level, and (d) a 10 ml. pipette properly filled with water.

EXPERIMENT 2. The determination of density.

1. Weigh an empty beaker. Pour a carefully measured volume of water into the beaker, using either a graduated cylinder or a pipette, and weigh the beaker and the water.

What is the density of water in grams per ml.?

Describe how you measured the volume of the water.

2. Use a lump of roll sulphur about one half inch in diameter, but small enough to permit dropping it into a graduated cylinder. Weigh the sulphur and record its weight.

What is the weight, in grams, of the lump of sulphur?

Add water to a graduated cylinder until it is approximately half full of water and record the volume. Drop the piece of sulphur into the water and record the volume.

What is the difference in volume and what does it represent?

What is the volume of the sulphur in cubic centimetres?

Find the density of sulphur in grams per cc. What is the specific gravity of sulphur?

3. Repeat the above experiment using a piece of iron instead of sulphur.

What is the density of iron in grams per cc.? What is the specific gravity of iron?

QUESTIONS

1. If the iron used in the above experiment formed a perfect cylinder, describe how you could find the density of iron without using the method of displacement.

2. Describe how you could alter Procedure 1 in Experiment 2 to get a more accurate value for the density of water.

EXPERIMENT 3. Changes in state.

1. Heat a test-tube half full of water to the boiling point. Place a small piece of Wood's metal in the boiling water and observe its behaviour. Cool the water and note any changes in the metal.

Describe the changes that take place in the metal. Are these physical or chemical changes? What names are used to express these changes in state?

Suggest how you could find the melting point of Wood's metal.

State an industrial use of Wood's metal that is a direct application of a property learned from this experiment.

2. Place some crystals of naphthalene in a clean, dry test-tube to a depth of one inch (*Technique 8*). Half fill a beaker with water and set it on a wire gauze placed on a ring attached to a retort stand. Put the test-tube containing the naphthalene into the water, and place a thermometer in the naphthalene so that the bulb is surrounded by the crystals. Heat the water until it becomes luke-warm and then lower the flame so that its temperature rises slowly. Observe the temperature at which the naphthalene melts. Continue heating until the naphthalene has completely liquefied, and then remove it from the beaker. Note the temperature at which solidification begins.

At what temperature did the naphthalene melt?

What was the temperature at which solidification began? What is the melting point of naphthalene? Did the temperature change during solidification?

What two changes in state are illustrated in this experiment?

3. Recall experiments that you have performed in previous grades, and answer the following questions.

Is it correct to say that the boiling point of water is $100^{\circ}\text{C}.$?

What condition must be recognized when reference is made to the boiling point of any liquid?

Define the boiling point of a liquid.

What term is used to express the change in state from a gas to a liquid? What is the heat relationship in such a change?

4. Place a spoonful of naphthalene in an evaporating dish (Pyrex is preferable). Cover the dish with a filter paper, and then invert another evaporating dish over the first one so that the filter paper separates the two dishes (Fig. 20). Heat the naphthalene gently until

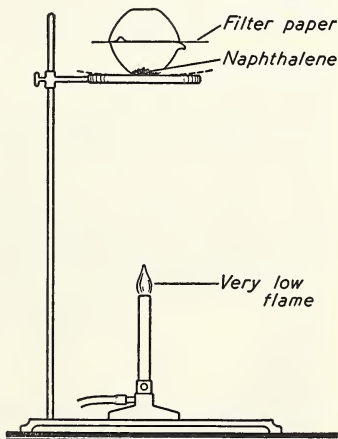


Fig. 20. The sublimation of naphthalene.

it has disappeared from the lower dish. (If it should melt during the heating, *reduce the temperature.*) Examine the filter paper and the upper evaporating dish.

What happened to the naphthalene in the lower dish?

Describe what was observed on the filter paper and upper dish.

What must have been the state of the naphthalene to allow it to pass through the filter paper?

What change in state is illustrated in this experiment? (TEXT p. 11)

QUESTIONS

1. Describe an experiment to show that if the pressure is reduced the boiling point of water is lowered.

2. Describe an experiment to show that if the pressure is increased the boiling point of water is increased.

3. Give an example from the above experiments for each of the changes in state shown in Figure 2.1 on page 9 of the text.

4. Give an example of a sublimation that occurs in nature, and describe the conditions under which such a phenomenon occurs.

EXPERIMENT 4. Physical and chemical properties.

1. Examine a lump of roll sulphur and a piece of iron such as were used in Experiment 2. Note the colour and odour of both substances. Wet a finger and touch it to the sulphur and then to the tongue. Repeat with the iron. Attempt to scratch each in turn with a knife. Attempt to break each substance into two pieces. Bring a magnet to touch each substance in turn. Shake a small piece of sulphur and a few coarse iron filings, separately, in about 5 ml. of water in a test-tube. Filter, collect a few drops of each filtrate on separate watch-glasses, and evaporate to dryness (*Technique 5*).

By completing the following table in your notebook, tabulate your observations so as to compare sulphur and iron in the following properties: colour, odour, taste, density (Experiment 2), hardness, brittleness, magnetic or non-magnetic property, and solubility in water.

PHYSICAL PROPERTIES	ROLL SULPHUR	IRON
Colour		
Odour		
Taste		
Density		
Hardness		
Brittleness		
Magnetic or non-magnetic		
Solubility in water		

Why is each of the above properties called a physical property?

2. Put a small piece of sulphur about the size of a pea on a piece of mica placed across a ring supported on a retort stand, and direct the hot flame of a Bunsen burner onto the sulphur. Observe any changes in colour, state, and odour. Repeat with iron.

Did the sulphur undergo any permanent change? Give reasons for your answer.

What is the effect of heat on the iron?

What property of sulphur and what property of iron is illustrated by this experiment? Are these properties physical or chemical? Explain.

EXPERIMENT 5. The identification of a few common substances.

1. Place samples of salt, sugar, starch, powdered zinc, and baking soda on separate pieces of paper, and number the papers according to a prearranged plan. The substances selected are harmless and can be touched, tasted, and smelled without danger, but the identification of unknown substances by these methods is not recommended as a general practice.

Record the colour, odour, and taste of each substance examined.

2. Put about 5 gm. of each substance, separately, on a piece of mica placed across a ring supported on a retort stand, and turn the hot flame of a Bunsen burner directly on each one.

Does the substance burn? Is any permanent change observed?

3. Into separate test-tubes, place a small quantity of each substance. Add a few drops of hydrochloric acid to each, and observe.

Describe any evidence of a chemical reaction between the acid and the substance tested.

4. Into separate test-tubes, place a small quantity of each substance. Add 5 ml. of distilled water to each one, and shake. Heat each one to the boiling point, and cool. Save these test-tubes and their contents for Procedure 5.

Are the substances soluble in cold water?

What was the effect of heat in each case?

5. Divide the liquid in each test-tube prepared in Procedure 4 into two equal parts. To one portion, add a few drops of silver nitrate solution, and to the other, a few drops of tincture of iodine.

Record the effect of each reagent on each substance.

QUESTIONS

1. Collect the information learned about each substance, and tabulate the properties of each under the headings: colour, odour, taste, solubility in water, flammability, effect of hydrochloric acid,

and the effect of adding (a) silver nitrate solution, and (b) tincture of iodine, to a solution of each substance (if the substance is soluble).

2. From a comparison of the properties tabulated above, tell how you could distinguish each substance from the other four. Give as many differences as you can.

EXPERIMENT 6. Some interesting chemical changes.

1. CAUTION: Do not inhale any gas produced.

Put a pellet of copper or a few copper turnings in a test-tube, and add a few drops of concentrated nitric acid (*Technique 2*).

List three observations.

What evidence is there that at least two substances, other than the original copper and nitric acid, are present?

2. Examine samples of lead nitrate and potassium iodide. To 5 ml. of a solution of lead nitrate in a test-tube, add 5 ml. of a potassium iodide solution.

Describe the two substances used.

From the fact that you used solutions of these substances, what can be said about their solubilities?

What happens when solutions of the two substances are mixed?

Give two observations.

Why would you conclude that at least one new substance has been formed?

3. On a piece of paper mix about 1 gm. of baking soda and 1 gm. of tartaric acid. Place the mixture in a test-tube, and then add about 5 ml. of water.

What evidence is there of a chemical reaction?

4. (*Demonstration*) Mix one volume of powdered sugar with two volumes of potassium chlorate, and place a half-teaspoonful of the mixture on a piece of mica supported by a ring attached to a retort stand. The base of the stand should be resting on an asbestos mat to prevent possible damage to the desk surface. Touch the mixture with a glass rod dipped in concentrated sulphuric acid.

What evidence is there of a chemical reaction?

EXPERIMENT 7. Heating metals in air.

1. Clean a 1-inch square of copper foil with emery paper, and carefully examine the cleaned surface. Using tongs, hold the foil in the tip of a hot Bunsen flame and continue to heat it until no further change is noticeable. Allow the foil to cool, and examine the surface again carefully. With a knife, scrape off some of this surface coating and examine the flakes closely.

Describe the colour, lustre, and flexibility of the copper foil before heating.

Describe the colour, lustre, and flexibility of the flakes scraped from the surface of the foil. Can you distinguish two colours?

Explain the nature of the change that has taken place.

Clean the foil with emery paper once more, examine the cleaned surface, and repeat the heating.

Why is it necessary to clean the foil for the chemical change to continue?

2. CAUTION: Care should be taken that burning magnesium is never held over an unprotected desk surface.

Clean a piece of magnesium ribbon, not more than three inches long, by drawing it through a piece of folded emery paper or fine sandpaper, and carefully examine the cleaned surface. Roll the cleaned ribbon into a compact coil, and **using tongs**, hold the tip of the ribbon in a Bunsen flame. When it ignites, hold the burning ribbon over an asbestos mat and catch the product of combustion. Examine the new substance formed.

List four properties of magnesium.

Describe the action that takes place when magnesium is heated.

Describe the substance formed by the burning of magnesium, and compare its properties with those of magnesium.

Has there been a physical or a chemical change? Give reasons for your answer.

3. Examine some clean steel wool. (Although not pure iron, steel wool closely resembles iron.) Using tongs, hold the steel wool in a hot Bunsen flame for a short time. Remove it from the flame, allow it to cool, and look for any new substance that has been formed.

Compare the properties of the new substance with those of iron.

What kind of change has taken place?

4. (Demonstration) Examine a clean platinum wire. Hold it in a hot Bunsen flame, and observe any change in appearance while it is in the flame. Allow it to cool, and examine the platinum again.

Is there any permanent change in the platinum?

QUESTIONS

1. Do all metals undergo a chemical change when they are heated in air?

2. Name two metals, other than those used in this experiment, that were heated in air by Lavoisier. (TEXT p. 20).

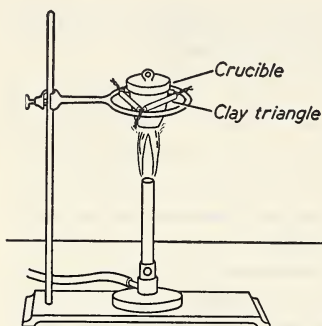
EXPERIMENT 8. Change in weight when a metal is heated in air.

Fig. 21. Heating a covered crucible and its contents.

1. Support a crucible with its cover on a clay triangle, which in turn is supported by a ring attached to a retort stand (Fig. 21). Heat the crucible and its cover, gently at first, and then, for at least one minute, with the hottest flame of your burner. Cool to room temperature, and then weigh the crucible together with its cover.

2. Roll a piece of clean magnesium ribbon, about four inches in length, into a loose coil, and place it in the crucible. Weigh the crucible, cover, and magnesium.

3. Return the crucible and its contents to the clay triangle, and place the *inverted* cover loosely over the crucible so that there is a narrow space through which air may enter. Heat, gently at first, and then more strongly, until all the magnesium has burned. It may be necessary (using tongs) to adjust the cover from time to time to admit air to the crucible, but this must be done with care or some of the ash may escape. Cool, and weigh the crucible, lid, and contents.

4. Record the weights as follows:

1. Weight of crucible and cover gm.
2. Weight of crucible, cover, and magnesium gm.
3. Weight of crucible, cover, and ash gm.

5. Calculate:

- (a) Weight of magnesium used (2 - 1) gm.
- (b) Weight of ash formed (3 - 1) gm.

How much did the magnesium increase in weight when it was burned? What is the cause of this increase in weight? What percentage is this of the original weight?

6. Some of the class may substitute copper for magnesium in the above experiment. If copper is used, very fine wire or foil is suitable, and the crucible containing the copper should be heated with a high flame for at least 15 minutes.

QUESTIONS

1. Account for the increase in weight experienced by some metals when they are heated in air.
2. Why was the crucible and its cover heated before it was weighed?

UNIT 2

Oxygen

EXPERIMENT 9. Heating mercuric oxide.

1. Place about as much mercuric oxide as will lie on a ten cent piece in a small Pyrex test-tube (*Technique 8*). Hold the test-tube as in Figure 22, and heat gently, keeping the end containing the solid in the flame. After heating for a very short time, remove the test-tube from the flame and allow it to cool.

List the changes in the appearance of the mercuric oxide when hot and when cold.

2. Continue to heat the mercuric oxide, increasing the temperature until it is being strongly heated. Lower a glowing wood splint into the test-tube slowly, until it almost touches the material in the bottom of the tube, and just as slowly, withdraw it. For a convenient reference, the accompanying diagram (Fig. 22) is numbered.

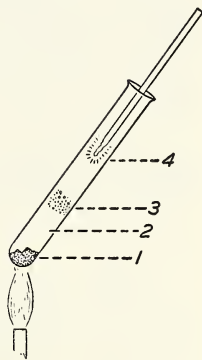


Fig. 22. Heating mercuric oxide.

Copy Figure 22 in your notebook.

Describe and account for what happened to the glowing splint when it was placed in Position 4.

Describe the change in the condition of the splint as it was lowered to Position 2.

Give two reasons for concluding that the mercuric oxide was undergoing a chemical change.

Describe the change in the splint when it was again raised to Position 4.

3. Continue heating the test-tube until the mercuric oxide in the bottom has disappeared. Remove the tube from the flame and allow it to cool. Examine the sides of the tube.

Describe the colour and lustre of the material found at Position 3. What is its physical state?

Name the substance present at Position 3.

Describe and identify the substance seen just below Position 3.

4. Reconsider the observations made when the glowing splint was raised and lowered in the test-tube while the mercuric oxide was being strongly heated.

Account for the change in the condition of the splint as it was moved from 4 to 2, and also for the change when it was moved from 2 to 4.

What is the factor that caused the change in the rate of burning of the splint?

Name the two substances that are present in the gaseous state at 2.

Which of these condenses at 3? Why?

Why does the other substance not condense?

A note to the teacher: To avoid damage to plumbing, students should be instructed to return the test-tubes used in this experiment to the teacher's desk. These can be treated with dilute nitric acid to form mercuric nitrate, and the mercury content recovered by displacement with zinc.

An alternative procedure is to have each student dislodge the mercury film by scraping with a wood splint, and then pour the mercury droplets into a special container provided by the teacher.

QUESTIONS

1. Write a word equation to represent the chemical reaction that takes place when mercuric oxide is heated.

2. Who was the first chemist to prepare oxygen from mercuric oxide? What was the name by which this chemist knew mercuric oxide? What was the name he used for oxygen? (TEXT p. 19)

3. When mercuric oxide is heated, **decomposition** takes place. When mercuric oxide is formed, the action is called **combination**. Compare the temperatures at 1 and 3 (Fig. 22), and draw a conclusion as to which is the lower temperature, the temperature of decomposition or the temperature of combination.

4. Would mercuric oxide be a suitable source of oxygen in quantity in the laboratory? Give reasons for your answer.

EXPERIMENT 10. The effect of heat on some substances that contain oxygen.

Put about one-half inch of each of the following substances (each substance listed contains oxygen) into separate Pyrex test-tubes: potassium permanganate, sand, calcium oxide, lead dioxide, potassium nitrate, sodium nitrate, cupric oxide, and potassium chlorate. Heat each test-tube, gently at first, and then more intensely. Test each one with a glowing splint.

Do all substances that contain oxygen liberate oxygen when they are heated?

From which of the above substances is oxygen liberated by heating?

List these substances in the order of the ease with which they liberate oxygen.

EXPERIMENT 11. Oxygen from potassium chlorate and the effect of manganese dioxide as a catalyst.

CAUTION: Do not use soft-glass test-tubes in this experiment.

Students are warned against allowing the carbon from wood splints to fall into molten potassium chlorate or a mixture of potassium chlorate and manganese dioxide. If this should occur, heating must be stopped at once. Accidents are less probable if long 8- to 10-inch pine or cedar splints are used instead of the shorter types obtained from supply houses.

1. Place crystals of potassium chlorate in a dry Pyrex test-tube to a depth of about one-half inch, and heat gently. Gradually increase the temperature until the potassium chlorate melts, and then bring a glowing splint to the mouth of the tube. Continue heating until bubbles begin to come off. Again bring the glowing splint to the mouth of the test-tube.

What is the cause of the crackling sound heard when crystals of potassium chlorate are first heated? (TEXT pp. 26 and 98)

Is there any evidence that oxygen is being given off from the potassium chlorate just at the time it is melting?

What gas is coming from the molten bubbling potassium chlorate?

Continue the heating, and test from time to time with a glowing splint until the splint will not rekindle. The substance left in the test-tube is called potassium chloride.

What change in state is noted in the substance in the test-tube?

Why did the bubbling cease?

Which has the lower melting point, potassium chlorate or potassium chloride? Give a reason for your answer.

2. Heat the same quantity of potassium chlorate as used in Procedure 1 until it just melts. Drop in a pinch of manganese dioxide, and test with a glowing splint.

Note any change in the liquid and the effect on a glowing splint held at the mouth of the tube.

Compare the temperature at which oxygen was liberated in Procedure 1 with that in Procedure 2. What are the effects of adding the manganese dioxide?

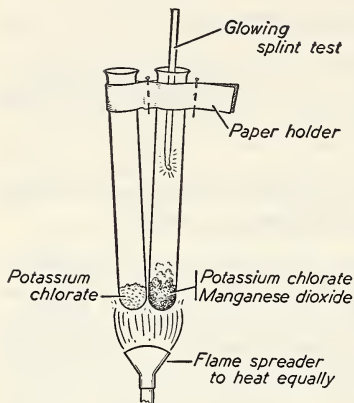


Fig. 23. The effect of heating manganese dioxide with potassium chlorate.

3. Mix together on a piece of paper approximately the same quantity of potassium chlorate as used in Procedure 1 and approximately one-third this quantity of manganese dioxide. Place this mixture in a Pyrex test-tube. In another Pyrex test-tube, place an equal quantity of potassium chlorate only. Hold both test-tubes in a common holder and, using a flame spreader, heat the two uniformly (Fig. 23). Test each tube at frequent intervals with a glowing splint.

Set aside the test-tube containing the manganese dioxide for use in Procedure 4.

From which tube is the oxygen liberated first?

From which is it liberated more rapidly?

What is the effect of having manganese dioxide present?

4. Add water to half fill the test-tube saved from Procedure 3, and heat until near boiling. Filter (**Technique 4**).

Describe the residue on the filter paper.

What substances are present in the filtrate?

5. Dry the filter paper containing the residue from Procedure 4, and scrape a small amount of the dried residue from the paper. Place some potassium chlorate crystals to a depth of about one-half inch in a Pyrex test-tube. Heat the tube until the potassium chlorate just melts, and then add a pinch of the dried residue. Test with a glowing splint.

What was observed when the residue was dropped into the molten potassium chlorate? What gas was liberated?

What property has this residue in common with manganese dioxide?

In what other respects does the residue resemble manganese dioxide?

QUESTIONS

1. Write a word equation expressing the reaction that takes place when potassium chlorate is strongly heated.

2. What name is given to express the type of reaction resulting from the use of manganese dioxide in the above experiment? (TEXT p. 27)

3. What conclusions can be made regarding the action of manganese dioxide on potassium chlorate in the production of oxygen?

4. Suggest an experiment designed to determine if any manganese dioxide has been used up in this reaction.

EXPERIMENT 12. (*Demonstration*) Weight relations when manganese dioxide is used as a catalyst with potassium chlorate in the preparation of oxygen.

1. Put about 1 gm. of manganese dioxide on a filter paper and accurately weigh the two together. Save the filter paper for future use. Mix the manganese dioxide with approximately three times its volume of potassium chlorate, and put the mixture in a Pyrex test-tube. Heat the mixture until all bubbling ceases and then continue to heat for about five minutes.

2. When the test-tube is cool, half fill it with distilled water, and bring the liquid nearly to the boiling point. Filter the liquid using the filter paper on which the manganese dioxide was originally weighed. To remove all of the material from the test-tube, add distilled water from time to time, each time pouring the liquid into the filter paper. Continue to wash the test-tube in this manner until all of the solid is removed. Wash the residue on the filter paper thoroughly with distilled water (*Technique 9*). Allow the residue and filter paper to dry completely.

3. Weigh the filter paper with the attached residue.

What is the residue?

Is there any change in the weight of the manganese dioxide?

What additional information does this experiment give about the action of the manganese dioxide?

Why was the residue so thoroughly washed?

EXPERIMENT 13. To prepare and collect oxygen and study its properties.

PREPARATION AND COLLECTION

1. Mix thoroughly, on a piece of paper, a teaspoonful of potassium chlorate with approximately one-third of a teaspoonful of manganese dioxide, and fill a Pyrex test-tube to one quarter of its depth with the mixture. Hold the test-tube in a nearly horizontal position and distribute the mixture along the sides of the lower half of the tube.

2. Set up the apparatus as shown in Figure 24. Before the gas bottles are placed in the pneumatic trough, fill each with water until

the rounded upper surface of the water extends above the mouth. Slide a glass plate over the mouth of the bottle and hold it in this position until the bottle is inverted in the water in the pneumatic trough, and then remove the glass plate (*Technique 12*).

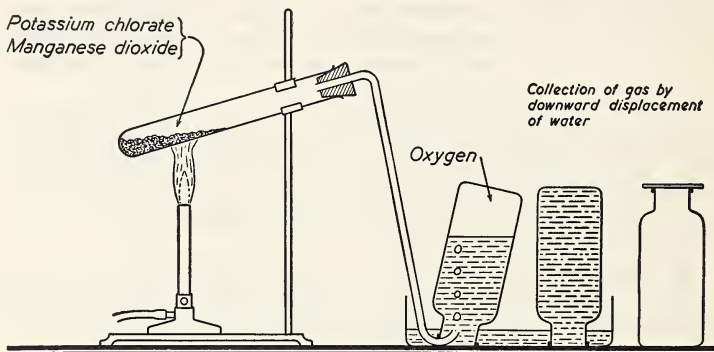


Fig. 24. The laboratory preparation and collection of oxygen.

3. Begin heating the mixture nearest the mouth of the tube and gradually work towards the base. Heat gently so as to obtain a slow but steady stream of bubbles in the pneumatic trough. Reject the first bottle of gas collected. *Why?*

4. When each bottle is full, cover the mouth with a glass plate and lift it out of the water, placing it mouth upward on the desk. When the last bottle has been collected, lift the delivery tube out of the water in the trough before removing the burner. *Why?*

A note to the teacher: More than one period will be required to complete the work that follows. It is suggested that the students might prepare and collect three bottles of oxygen at each desk. One bottle may be used for the glowing splint test and the identification of the resulting oxide with limewater. The second may be used for the formation and identification of an acidic oxide, using minute masses of roll sulphur. The third might be used for the formation and identification of a basic oxide, using coiled magnesium ribbon.

It is suggested that the remaining experiments be *demonstration experiments* in a succeeding lesson or lessons. The oxygen necessary for these demonstration experiments may be obtained from cylinders available from many supply houses.

PROPERTIES OF OXYGEN

1. Examine a bottle of oxygen.

Note the colour and odour of the gas. What evidence have you that oxygen is not very soluble in water?

2. Insert a glowing splint into a bottle of oxygen. Extinguish the flame and repeat the experiment several times, using the same bottle.

Compare the rate of reaction in oxygen with the rate in air.

Does oxygen support combustion?

Does oxygen burn?

3. Pour approximately 10 ml. of limewater into a bottle of oxygen, and shake for half a minute.

Is there any change in the appearance of the limewater?

Heat a piece of wood charcoal in a deflagrating spoon until it glows, and then lower it quickly into the bottle of oxygen containing the limewater. Allow the charcoal to remain there until no further action takes place, and then remove it.

Compare the burning of charcoal in air and in oxygen.

What happens to the limewater?

What gas was formed in the bottle? (TEXT p. 34)

Write a word equation for the burning of charcoal.

4. Lower a piece of glowing charcoal into another bottle of oxygen, and hold it there until no further action takes place. Then remove it. Pour about 10 ml. of water into the bottle, and shake vigorously for half a minute. Pour half of this liquid into a beaker.

Place a small piece of red litmus paper and a small piece of blue litmus paper in the liquid remaining in the bottle, and leave for several minutes. If you are in doubt as to whether or not there has been any change, put small pieces of the two litmus papers into distilled water, which does not affect litmus, and compare them with the colour of the litmus in the liquid being tested.

Describe any change in the colour of either of the litmus papers.

Add a few drops of neutral (green) bromthymol blue solution to the liquid in the beaker.

Note any change in colour.

Which is the better indicator in this experiment, litmus or bromthymol blue? Why?

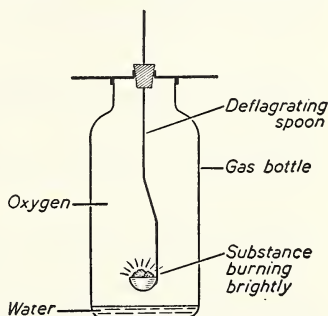


Fig. 25. Burning substances in oxygen.

What property of the solution is shown by the use of these indicators? (TEXT p. 34)

5. Line a deflagrating spoon with asbestos paper, and in it, place a piece of roll sulphur about the size of a bean. Ignite the sulphur in a Bunsen flame.

Note the colour, size, and brightness of the flame and any change in state or colour of the sulphur.

Lower the burning sulphur into a bottle of oxygen.

Note any changes in the flame.

After the flame is extinguished, remove the spoon from the bottle. *Is there any unburned sulphur remaining in the spoon?*

Why did the flame go out?

Smell the gas remaining in the bottle, using your hand to waft a little of the gas towards your nose (*Technique 14*).

Describe the odour.

What is the name of this gas? (TEXT p. 34)

Write a word equation for the burning of sulphur in oxygen.

Add about 10 ml. of water to the bottle, shake for half a minute, and test the liquid with litmus papers (as in Procedure 4).

What do the results indicate about the nature of this liquid?

Name this solution. (TEXT p. 34)

Write the word equation that represents the formation of this new substance.

6. Wind a piece of magnesium ribbon, about three inches long, into a compact spiral by twisting it around a match or a piece of glass tubing. Remove the spiralled ribbon and hold it in a pair of tongs. Place a protective asbestos mat under a lighted Bunsen burner, and ignite the tip of the spiral in the flame.

What is observed just before the magnesium ignites?

Lower the burning magnesium into a bottle of oxygen, also on a protective mat, being careful not to let it contact the sides of the bottle.

Compare the burning of magnesium in air and in oxygen.

Describe the nature of the oxide found in the bottle, and name it.

Add about 10 ml. of water to the bottle, shake, and test the liquid with red and blue litmus papers.

Does any of the material dissolve?

What do your observations indicate about the nature of this liquid?

Name the substance present in the solution. (TEXT p. 34)

Write a word equation representing the action that takes place when the product of combustion is shaken with water.

7. (*Demonstration*) Place a pinch of red phosphorus in an asbestos-

lined deflagrating spoon. Ignite it and quickly introduce the spoon into a bottle of oxygen.

Compare the burning of phosphorus in air and in oxygen.

Why is the burning of phosphorus more rapid in oxygen than in air?

Describe the product of the combustion of phosphorus in oxygen.

Phosphorus forms two oxides (TEXT p. 370). When there is a high concentration of oxygen, phosphorus pentoxide is formed. As the oxygen concentration diminishes, phosphorus trioxide is produced. In this experiment, the product of combustion is probably a mixture of these two oxides. *Why?*

Add about 10 ml. of water to the bottle, shake, and test with litmus.

Classify the liquid as acidic or basic.

Write two word equations for the burning of phosphorus in oxygen.

Write two word equations to show the reactions when the products of combustion are shaken with water.

A note to the teacher: If white (yellow) phosphorus is used, the result is more spectacular, but great care is necessary. This form of phosphorus is kept under water and must be cut under water. It must be handled with forceps: never with the hands. Any unburned phosphorus remaining on the spoon should be destroyed by complete burning or by immersing the spoon in cupric sulphate solution. Even a trace of white phosphorus exposed to the air could be the source of a major fire.

8. (Demonstration) Place a small piece of sodium in an asbestos-lined deflagrating spoon, and heat it until it ignites.

Describe any changes in the sodium before it ignites.

Does it ignite easily?

Compare the ignition temperatures of sodium and phosphorus.

What is the colour of the flame?

Lower the burning sodium into a bottle of oxygen.

Describe the burning in oxygen and the product of combustion.

Could the product contain a mixture of oxides? (TEXT p. 349)

Shake the product of combustion with water, and test with litmus.

Write word equations to represent these reactions.

9. (Demonstration) Add sufficient water to a bottle containing oxygen to occupy a depth of one inch. Using tongs, heat a piece of fine iron wire (or steel wool) until it glows brightly. Quickly lower it into the bottle of oxygen.

Record your observations.

Why was water added to the bottle before the introduction of the iron?

Compare the colour and brittleness of the product of combustion (magnetic iron oxide) with that of the original iron.

Shake some of the oxide with water, and test with litmus.

Does magnetic iron oxide appear to be soluble in water? (TEXT p. 34)

10. To a bottle filled with oxygen, add 5 ml. of a solution of pyrogalllic acid and 5 ml. of a solution of potassium hydroxide. Place the hand firmly over the mouth of the bottle and shake.

What is the effect of oxygen on a mixture of these two solutions?

What effect was noticed on the hand? What does this indicate?

The above reaction is used as a conclusive test for oxygen. (TEXT p. 33)

QUESTIONS

1. Make a neatly labelled diagram of the apparatus used to prepare and collect oxygen.

2. Name three methods used in the collection of a gas. (TEXT p. 28) Which method was selected for the collection of oxygen? Why?

3. Summarize the results of this experiment by completing the following table in your notebook:

ELEMENT BURNED	METAL OR NON-METAL	COLOUR OF FLAME	DESCRIPTION OF OXIDE	NAME OF OXIDE	EFFECT OF AQUEOUS SOLUTION ON LITMUS	NAME OF ACID OR BASE FORMED	CLASS OF OXIDE
Carbon							
Sulphur							
Phosphorus				{ 1. 2.		{ 1. 2.	
Magnesium							
Sodium				{ 1. 2.		{ 1. 2.	
Iron							

4. Name, define, and give an example of each of the three classes of oxides.

5. What type of element gives rise to (a) an acidic oxide, (b) a basic oxide?

6. What is produced when an acidic oxide reacts with water? Give an example.

7. What is produced when a basic oxide reacts with water? Give an example.

EXPERIMENT 14. The effect of adding water to sodium peroxide.

CAUTION: Care must be taken to avoid contact between sodium peroxide and moist skin.

If paper or any other easily combustible substance is used for pouring the sodium peroxide into the test-tube, it should not be

discarded until it has been placed in the sink and thoroughly soaked with water.

1. Clamp a test-tube to a stand and add sodium peroxide to the tube to a depth of one-half inch. Add a few drops of water, and feel the bottom of the tube. Test at the mouth of the tube with a glowing splint.

Is the reaction between sodium peroxide and water endothermic or exothermic in nature?

What gas is produced by this reaction?

2. Add enough water to half fill the test-tube, and when all action ceases, test the solution with neutral litmus paper.

What is a second product of the reaction between sodium peroxide and water?

Write a word equation for the reaction between sodium peroxide and water?

Would sodium peroxide be a suitable source of oxygen in quantity?

EXPERIMENT 15. (Demonstration) Ignition or kindling temperatures.

1. **CAUTION:** Do not touch the phosphorus with the hands (*Technique 10*). Procedure 1 should not be performed by students.

Using filter paper, dry a piece of white (yellow) phosphorus about the size of a grain of wheat, and place it on a piece of mica or asbestos paper. Put a piece of sulphur of approximately the same size near it. Heat a glass rod, not warmer than can be held with the hand, and touch it to the sulphur and then to the phosphorus.

Describe what happens.

Which has the lower ignition temperature, sulphur or phosphorus?

Why is the precaution given about handling white (yellow) phosphorus?

2. Examine a "strike anywhere" match. The tip is phosphorus sesquisulphide, a substance with a very low ignition temperature. The material surrounding the tip is chiefly potassium chlorate coloured with a suitable dye. Rub the match on a piece of fine sandpaper.

Make a diagram of the match and label all the parts.

Describe and explain what happens when the match is struck on the sandpaper. What is the purpose of (a) the phosphorus sesquisulphide, and (b) the potassium chlorate?

Note any change in appearance in the wood near the flame. Explain.

What two effects has an increased temperature on paraffin?

The wood of a match has been dipped in a water solution of alum or sodium sulphate or some other suitable chemical, and then dried.

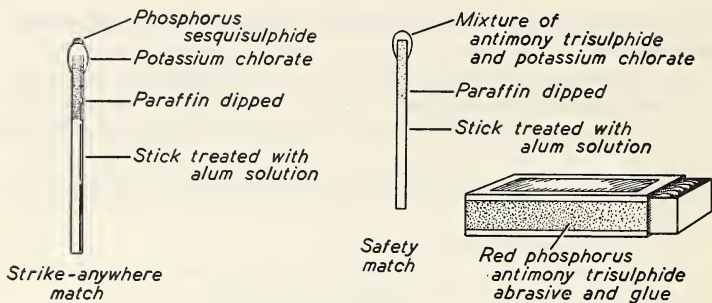


Fig. 26. "Strike anywhere" and "safety" matches.

The wood is so treated to prevent smoldering. Light a strike-anywhere match and allow it to burn about half way down. Blow it out.

Why does blowing extinguish the flame?

Is there an after-glow?

Ignite a wooden splint. Allow it to burn for a short time and then blow it out.

Is there an after-glow? Explain.

3. The tip of a "safety" match usually consists of a mixture of antimony trisulphide, potassium chlorate, and glue. The material on the side of the match container (packet or box) is a mixture of red phosphorus, antimony trisulphide and powdered glass, held together with glue. Rub a safety match on a piece of fine sandpaper, and then rub it on the material provided for striking.

Make a diagram of a safety match and label all the parts.

What happens when the match is rubbed (a) on sandpaper, and (b) on the special material provided?

Friction changes a trace of the red phosphorus to the white (yellow) allotropic form (TEXT p. 368), which has a much lower ignition temperature than the red form.

Explain why a safety match will ignite on the relatively smooth surface provided for striking.

Is there any after-glow in the safety match?

4. (**Demonstration**) Put one-half a teaspoonful of corn starch on a piece of mica placed across a ring supported on a retort stand, and turn a Bunsen flame on it for some time. When the starch starts to burn, remove the flame.

Does the starch continue to burn?

Has starch a low or high ignition temperature?

CAUTION: Great care must be taken in the following experiment. The starch must be blown out of the paper cylinder with one quick, forceful puff, and the students should not be allowed within fifteen feet of the direction of blowing.

Put another half teaspoonful of starch in a paper cylinder made by rolling a piece of foolscap to about one-half inch in diameter. Hold one end of the paper tube about four inches from a Bunsen flame, and blow the starch from the tube into the flame.

Describe and explain what happens.

In comparing the ignition temperatures of two substances, why is it necessary to have conditions alike in all respects?

5. (*Demonstration*) Punch a hole in an eight-pound honey pail about one inch from the bottom. Bend a piece of metal tubing about five inches long to form a right angle, and solder a small metal funnel on one end. Push the metal tubing through the hole in the pail and solder it in position, as shown in Figure 27. Place a candle so that the wick is about an inch above the top of the funnel, and attach a piece of rubber tubing about two feet long to the metal tubing.

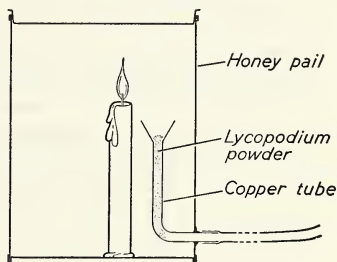


Fig. 27. A dust explosion.

Put a half-teaspoonful of lycopodium powder in the funnel. Light the candle, and leaving the cover off the pail, give a short quick puff into the open end of the rubber tubing, and observe.

Repeat the experiment but this time, after lighting the candle, push the cover on the pail and immediately give a quick puff into the open end of the tube, and observe.

Explain why a flash resulted when the pail was not covered, and an explosion occurred when the cover was on the pail.

QUESTIONS

1. Arrange the following substances in the order of their ignition temperatures: phosphorus sesquisulphide, wood, and paraffin.

2. Why does phosphorus sesquisulphide ignite when rubbed over a rough surface?

3. Why is white (yellow) phosphorus not used in matches?

4. Which has the higher ignition temperature: phosphorus sesquisulphide or antimony trisulphide? Give a reason for your answer.

5. Why does a large wooden stick not take fire as quickly as a small one?
6. Explain why dust explosions sometimes occur in grain elevators.

EXPERIMENT 16. (*Demonstration*) Spontaneous combustion.

CAUTION: This experiment should not be performed by students.

1. In a large test-tube containing approximately 5 ml. of carbon disulphide, dissolve a piece of white phosphorus about the size of a small pea. Pour the solution onto a piece of filter paper placed in the sink. Using tongs, remove the filter paper from the sink, place it on a wire gauze supported by a ring clamped to a retort stand, and let it stand for a few minutes.

Describe and account for the phenomenon observed.

A note to the teacher: Great care must be exercised in preparing and using this solution of phosphorus in carbon disulphide. If too much solution is made to be used at one time, it may be temporarily stored in a small, glass-stoppered bottle, provided a layer of water is added to cover the carbon disulphide solution. This bottle should be stored in a capped metal container.

To use the solution, squeeze the bulb of a medicine dropper and insert the glass tip below the water layer into the disulphide solution. The impregnation of the filter paper with this solution should be done over the desk sink. After use, the medicine dropper is contaminated with phosphorus and constitutes a fire hazard. It should be rinsed thoroughly and then stored in the same metal container as the bottle of solution.

When this topic has been completed with all classes, any remaining solution should be discarded, and all articles having traces of phosphorus on their surfaces made harmless by prolonged immersion in a solution of cupric sulphate.

2. (*Demonstration*) Place a ball of absorbent cotton about two inches in diameter on a wire gauze supported on a ring stand. Flatten the top of the cotton, and spread a teaspoonful of sodium peroxide in a layer over it. Add a few drops of water to the centre of the sodium peroxide.

CAUTION: The absorbent cotton used in this experiment should not be discarded until it has been thoroughly soaked with water to render it harmless.

Describe and account for the phenomenon observed.

Why should sodium peroxide never be left unguarded in the laboratory or in an uncovered bottle?

UNIT 3

Air

EXPERIMENT 17. (*Demonstration*) The preparation of impure nitrogen from air.

1. Hollow out a slight depression in the centre of a flat cork. Coat this surface with a mixture of plaster of Paris and water, and allow the mixture to harden. Float the cork on the surface of water in a pneumatic trough or the sink. The water should be about one inch in depth.

2. **CAUTION:** Do not touch the phosphorus with the hands (*Technique 10*). This experiment should not be performed by students.

Place a small piece of white phosphorus about the size of a grain of wheat in the depression in the cork. Warm a piece of iron wire in a Bunsen flame, and touch the phosphorus with the hot end of the wire. Quickly cover the cork and burning phosphorus with a wide-mouth bottle in such a way as to seal the mouth of the bottle beneath the water (Fig. 28). Keep the bottle in place until all action ceases.

Describe what is observed.

Account for all the changes that take place inside the bottle.

Why does the phosphorus stop burning before it is all consumed?

Name the gas left in the bottle and account for its presence there.

3. Place a glass plate over the mouth of the bottle. Remove the bottle and place it in an upright position on the table. Remove the glass plate and quickly insert a blazing splint. Insert another blazing splint in a bottle containing air.

Compare the behaviour of the burning splint in nitrogen and in air.

A note to the teacher: Any unburned phosphorus remaining on the cork should be destroyed by complete burning.

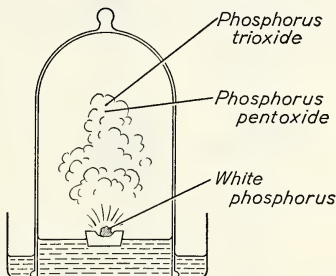


Fig. 28. Using white (yellow) phosphorus to prepare "atmospheric nitrogen".

QUESTIONS

1. List four impurities in the nitrogen obtained by this method. (TEXT pp. 43 and 47) What are the approximate percentages of the impurities?

2. Suggest how the above experiment could be modified so that the proportion by volume of oxygen and nitrogen in air could be determined.

EXPERIMENT 18. To find the proportions by volume of nitrogen and oxygen in air (using steel wool).

1. Degrease some steel wool by dipping it several times in gasoline. After the gasoline has evaporated, moisten the steel wool with water and push it into a large-mouth bottle (Fig. 29). Invert the bottle in a dish of water and allow it to stand for several days. Record observations from time to time.

2. After the water stops rising in the bottle, adjust the level of the water in the dish until it coincides with the level of the water in the bottle. When this position is attained, place a cover glass over the mouth of the bottle, and remove the bottle and its water contents from the dish.

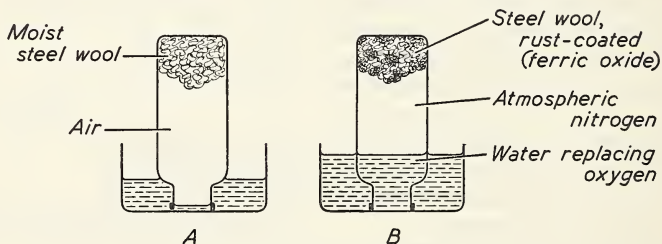


Fig. 29. The rusting of iron in air.

3. Use a graduated cylinder to measure the volume of the water in the bottle.

Record the volume of water that came into the bottle. What gas and what gas volume does this represent?

Fill the bottle with water, and measure the volume of water it contains.

Record the volume of water needed to fill the bottle.

What gas and what gas volume does this represent?

Describe any change observed in the steel wool. What is the significance of the fact that all the iron has not rusted?

Calculate the percentage by volume of nitrogen and of oxygen in air.

QUESTIONS

1. Why was the water level adjustment made before removing the bottle from the dish?

2. Point out all the possible sources of error in the above experiment.

3. Compare your calculation of the percentage of nitrogen in the air with the percentage shown in the TEXT, p. 47. What is your percentage error? (See page 3 of this manual.)

EXPERIMENT 19. (*Demonstration*) To find the proportions by volume of oxygen and nitrogen in air (using pyrogallic acid and potassium hydroxide).

1. Pour approximately 60 ml. of a concentrated pyrogallic acid solution into a 200 millilitre gas-measuring tube (Fig. 30). Slant the tube, and slide a stick of potassium hydroxide about an inch long (or a dozen small pellets) into it. Quickly insert a stopper, and holding the stopper in place, invert the tube.

Record the volume of air enclosed in the tube.

2. Shake the mixture slowly back and forth in the tube for about ten minutes.

What changes take place in the mixture?

3. Invert the tube in a deep vessel of water, and with the stoppered end well below the surface, remove the stopper.

Describe and account for what happens.

What gas remains in the tube?

4. Cool to room temperature, and adjust the level of the liquid inside the tube to coincide with the level of the water in the vessel.

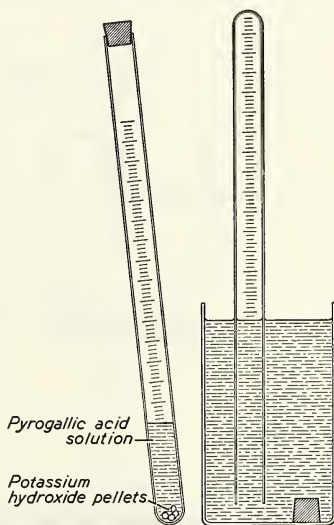


Fig. 30. Determining the composition of air by volume.

Record the volume of gas remaining in the tube. How could this gas be identified?

Why was the adjustment in levels made before recording this volume?

Calculate the percentage by volume of oxygen and of nitrogen in air.

What is your percentage error?

EXPERIMENT 20. (*Demonstration*) To prepare and collect pure nitrogen and study its properties.

1. Prepare a solution of sodium nitrite by stirring 20 gm. of the solid in 100 ml. of water. Place 30 ml. of a saturated solution of ammonium chloride in a flask fitted with a dropping funnel and a delivery tube (Fig. 31). Place the flask in a water bath, and *with the tap closed*, pour some of the sodium nitrite solution into the dropping funnel. Have a beaker of cold water handy.

CAUTION: A violent explosion may occur in the generator if the reaction becomes too vigorous. This experiment should not be performed by students.

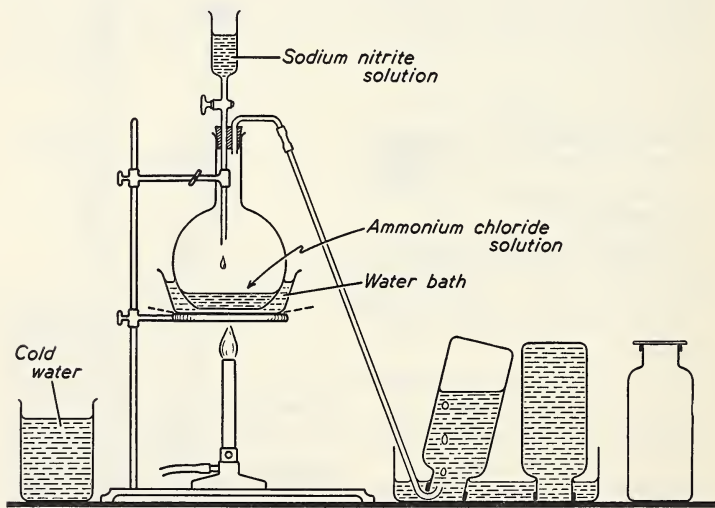


Fig. 31. The laboratory preparation and collection of pure nitrogen.

2. Heat the water in the water bath until warm, and begin adding the solution of sodium nitrite from the dropping funnel, drop by drop.

Should the contents of the flask froth considerably, remove the Bunsen flame and add cold water to the water bath so that the temperature is quickly reduced.

Collect three bottles of gas by the downward displacement of water. Discard the first bottle collected. *Why?*

3. Place the remaining two bottles in an upright position on the table, using a glass plate to retain the nitrogen. Insert a burning splint into one of the bottles of nitrogen, and burning sulphur in a deflagrating spoon into the other.

List four physical properties of nitrogen.

What two chemical properties of nitrogen are indicated by using the burning splint and burning sulphur?

QUESTIONS

1. Write two word equations for the preparation of pure nitrogen. (TEXT p. 44)

2. Why is it not advisable to prepare pure nitrogen by heating ammonium nitrite directly?

UNIT 4

Hydrogen

EXPERIMENT 21. The action of sodium and potassium on water.

1. (*Demonstration*) Using forceps, remove a piece of sodium from the container in which it has been stored, and place it on a *dry* glass plate.

What is the liquid in which sodium is stored?

Why is sodium stored in this liquid?

Describe the appearance of the sodium.

2. (*Demonstration*) Using a sharp knife, cut off a slice of sodium approximately one-eighth of an inch thick. Cut off any incrustation and discard these fragments, a little at a time, by placing them on the water in a partially filled sink.

What is the appearance of the freshly cut surface of sodium?

How would you describe the hardness of this metal?

Describe the changes in colour and lustre that take place on the surface of the sodium following exposure to air. Account for these changes.

A note to the teacher: Procedures 1 and 2 should not be done by students. The teacher should not permit students to cut the one-eighth inch cubes that are to be used in the class experiment suggested in Procedure 3. These should be cut by the teacher during or in advance of the lesson and stored under kerosene or mineral oil in a small, wide-mouthed bottle.

3. Using dry forceps, place a small sodium cube on a piece of filter paper. Using another piece of filter paper, gently pat the cube to remove any adhering liquid. Again pick up the sodium with forceps and drop it into a beaker half filled with water. **Immediately** cover the beaker with a glass plate or a larger inverted beaker, as illustrated in Figure 32, in order to protect the eyes should the sodium explode. If the sodium should stick to the glass, it should be dislodged by a gentle shaking of the beaker. Observe the behaviour of this active metal.

After all action has ceased, remove the protective cover, and add a piece of neutral (purple) litmus paper to the solution remaining in the beaker. Then dip the index finger of one hand into the solution and rub it against the thumb.

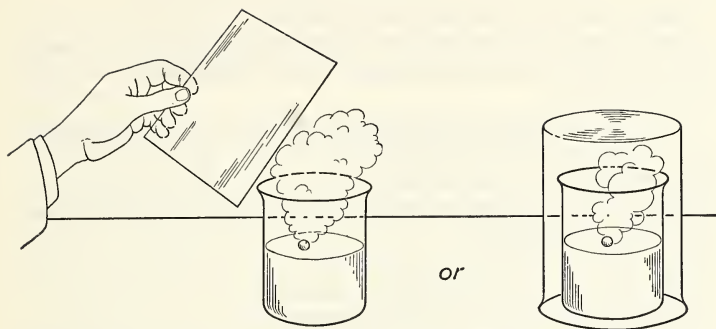


Fig. 32. Using a glass plate or inverted beaker in watching the action of sodium or potassium on water.

What change takes place in the shape of the sodium when placed on water? What is the position of the sodium relative to the surface of the water? Describe the motion of the sodium, and account for the sound produced. Describe the substance formed where the sodium is in contact with air. Describe the substance formed where the sodium is in contact with water. What evidence is there that the action of sodium on water is exothermic? What happens if the sodium clings to the side of the beaker? Describe the colour change in the litmus.

Describe the "feel" of the solution remaining in the beaker. What is the solute?

4. Wrap a cube of dry sodium loosely in a square inch of lead foil, and pierce the foil with the point of a knife blade several times. Drop the weighted sodium into a beaker of water, and observe the action.

Describe the action that takes place.

Why is it necessary to wrap the sodium in the foil?

5. Fill two test-tubes with water, and invert them in a beaker two-thirds filled with water. Repeat Procedure 4, but as soon as the weighted sodium has been dropped into the beaker, place one of the test-tubes over it and collect a test-tube full of the gas. Then quickly place the second tube over the weighted sodium and allow the tube to remain in this position until it is one-quarter filled with the gas. Keeping the tubes in an inverted position, remove both tubes from the beaker, and allow the water in the partly filled tube to be replaced by air. Keep the tubes in an inverted position and bring a blazing splint to the mouth of each test-tube in turn. Observe the action that takes place in both tubes.

What is the colour of the gas in the test-tubes?

Describe the action that takes place when the lighted splint is brought (a) to the test-tube full of the gas, and (b) to the test-tube in which the gas is mixed with air.

(The gas produced by the action of sodium on water is hydrogen, and the burning of the gas, as shown in 5(a) and 5(b), constitutes characteristic tests for hydrogen.)

6. Add a piece of neutral (purple) litmus paper to the solution remaining in the beaker.

Describe the colour change in the litmus.

What substance must be present in solution?

Account for the production of this substance.

7. Evaporate to dryness, in a test-tube, about 1 ml. of the solution remaining in the beaker.

What is the appearance of the residue in the test-tube?

What is the name of the residue? (TEXT p. 59)

8. (**Demonstration**) Half fill an evaporating dish or small beaker with water. Fold a small piece of filter paper so as to form a boat-like pocket, and float this folded paper on the surface of the water in the container. Using forceps, drop a small sodium cube into the water contained in the pocket.

CAUTION: The sodium may explode. This experiment should not be performed by students.

Describe the action of the sodium.

Account for (a) the flame, (b) the flame colour.

9. (**Demonstration**) Using potassium instead of sodium, repeat Procedures 1 to 7, but omit Procedure 8.

CAUTION: The reaction of potassium on water is more dangerous than sodium. When performing this experiment, great care must be taken not to have the face too near the globule of potassium. When the flame seems to disappear, the molten globule explodes and showers the immediate area with molten metal and strong alkali. Very tiny cubes of potassium should be used to minimize this dangerous spattering.

Observe the behaviour of this active metal and record your observations.

QUESTIONS

1. In what respects are sodium and potassium similar?
2. In what respects do sodium and potassium differ?
3. Which element is more active? Give reasons for your answer.
4. What is the source of the hydrogen obtained in the above experiments?

5. Write word equations to represent the action of (a) sodium on water, (b) potassium on water.

6. List the properties of hydrogen discovered as a result of these experiments.

EXPERIMENT 22. The action of calcium on water.

1. Half fill a 250 millilitre beaker with water, and invert four test-tubes filled with water into the beaker. Examine a small piece of calcium. Drop the calcium into the water in the beaker. Place the test-tubes in succession over the reacting calcium and collect (a) two test-tubes full of gas, (b) one test-tube half-full, and (c) one test-tube one-third full. Lift each test-tube from the beaker and stand it, mouth downward, on the table. Save the contents of the beaker for Procedure 3.

Describe the appearance of the calcium.

Compare calcium with sodium (Experiment 21) as to (a) hardness, (b) lustre, (c) density, and (d) activity with water.

Describe the appearance of the contents of the beaker.

Is there any evidence of an exothermic reaction?

2. Leaving one test-tube full of the gas on the table, lift the others, one at a time, and keeping the mouth down, bring a lighted splint to the open end of each tube. Hold the remaining test-tube full of gas, mouth upward, for one minute, and then bring a lighted splint to the mouth.

What characteristic property enables you to identify the gas in each case?

What is the reason for keeping the test-tube mouth down?

What happens to the gas in the test-tube held mouth up? Explain.

3. Add a piece of neutral litmus paper to the contents of the beaker (Procedure 1), and then filter some of the mixture (**Technique 4**). Collect approximately one-half of a test-tube full of the filtrate in one test-tube, and one-quarter of a test-tube full in another. Place one end of a clean glass tube under the liquid in the test-tube half filled with the filtrate, and blow your breath through the glass tube for several minutes. With a low flame, heat the tube one-quarter full of the filtrate, until no liquid remains in the tube (**Technique 7**).

Describe and explain the colour change of the neutral litmus.

What substance must be in solution to cause this colour change in litmus?

Account for the milky appearance of the contents of the beaker.

Describe the changes that take place when the breath is blown through the filtrate.

What is the common name for this solution?

Describe the appearance of the solid dissolved in the filtrate.

What is the chemical name of this solid?

4. Heat about 50 ml. of water in a beaker until it boils. Remove the flame, and when the water stops boiling, drop a small piece of calcium into the hot water.

How does the rate of this reaction compare with the rate of the reaction of calcium on cold water?

A note to the teacher: The authors recommend the use of British Drug Houses (B.D.H.) calcium in this experiment. Their mailing address is British Drug Houses (Canada) Limited, The Queensway, Toronto, Ontario.

QUESTIONS

1. In the above experiments, what evidence indicates that the hydrogen comes from the water?

2. From your observations, arrange the elements hydrogen, sodium, potassium, and calcium in the order of their relative activities. Do your experiments confirm the order as given in the TEXT p. 60?

3. Write a word equation to express the action of calcium on water.

EXPERIMENT 23. The action of magnesium on hot water and steam.

1. Half fill two beakers with water, and heat one to the boiling point. Into each, place a piece of clean magnesium ribbon and a piece of neutral litmus paper.

2. Repeat Procedure 1 using magnesium powder.

Describe the action of the magnesium ribbon on (a) cold water, (b) hot water.

Describe the action of the magnesium powder on (a) cold water, (b) hot water.

Do any of these experiments warrant collecting and testing the gas produced?

What conclusion can be drawn from the litmus test?

What effect was produced on the rate of reaction by (a) powdering the magnesium, (b) increasing the temperature?

3. (**Demonstration**) Half fill a 500 millilitre flask with tap water, and boil long enough to expel the dissolved air. Wind a piece of magnesium ribbon approximately six inches long into a tight spiral small enough to be lowered easily into the mouth of the flask. Attach the spiral to the bowl of a deflagrating spoon or pick it up with suitable

forceps or tongs, and after igniting the end, lower it carefully into the steam-filled flask. Closely observe the reaction that takes place.

What do you observe about the magnesium the instant before it starts to burn in air?

Does the magnesium continue to burn when surrounded by steam?

Is there any evidence of a gas burning just above the mouth of the flask? If so, describe the flame.

4. After the magnesium stops burning, keep a small amount of the ash from falling into the flask, and compare its appearance with the ash formed when magnesium burns in air. Add a piece of neutral litmus paper to the mixture of water and ash in the flask. Filter a small amount of the mixture in the flask, and collect some of the filtrate in a test-tube. Evaporate the filtrate to dryness.

How does the product formed by the burning of magnesium in steam compare with the product formed by the burning of magnesium in air?

What does the litmus test indicate concerning the solution?

What is the chemical name of the ash produced in this experiment?

Describe the residue remaining in the test-tube. What is its chemical name?

Write word equations to represent (a) the action of magnesium on steam, (b) the action of magnesium oxide on water.

QUESTIONS

1. What evidence is there to prove that hydrogen is produced when magnesium acts on steam?

2. What evidence is there to prove that magnesium oxide is produced when magnesium acts on steam?

3. Name two factors that affect the rate of a chemical reaction.

4. Compare the reaction rate of magnesium with (a) water at $100^{\circ}\text{C}.$, (b) steam at $100^{\circ}\text{C}.$ What is the reason for the difference?

5. As a result of this experiment place magnesium in its proper position in the activity series.

6. Suggest a method of collecting the hydrogen produced in an experiment using magnesium and steam.

7. Write word equations to represent (a) the action of magnesium on steam, (b) the action of magnesium oxide on water.

EXPERIMENT 24. (*Demonstration*) The preparation of hydrogen by the action of iron on steam.

1. Arrange the apparatus as shown in Figure 33, and heat the finely divided iron as intensely as possible while the steam is passing through

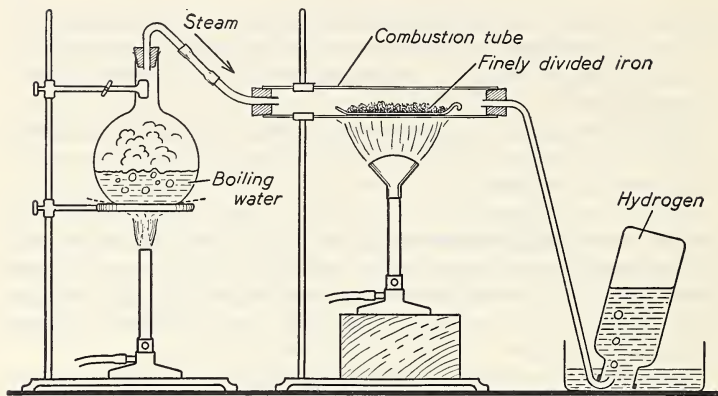


Fig. 33. Liberating hydrogen by the action of heated iron on steam.

the combustion tube. Collect a gas bottle full of gas, and test with a blazing splint.

Is there any evident chemical reaction within the combustion tube?

Is the gas produced hydrogen? Explain.

Is there any change in the appearance of the iron?

2. To compare the action of iron on steam with that of magnesium on steam, replace the porcelain boat containing iron with one containing small pieces of magnesium ribbon, and collect and test the gas produced.

How does the activity of iron on steam compare with the activity of magnesium on steam?

Using evidence gained from this experiment, add iron to the activity series, placing it in its proper position.

QUESTIONS

1. Write the word equation to represent the action of iron on steam.
2. Why is this method of producing hydrogen used to a great extent in the commercial preparation of hydrogen?

EXPERIMENT 25. The action on dilute acids of metals located above hydrogen in the activity series.

1. Into five separate test-tubes, place (a) a small piece of calcium, (b) some magnesium ribbon, (c) a small piece of mossy zinc, (d) an iron tack, and (e) a small amount of iron filings. Add a small amount (4 ml.) of dilute sulphuric acid to each test-tube in turn.

Note the reaction, and test the escaping gas with a lighted splint. Record all observations.

Is a gas given off in each case? Where a gas is given off, identify it.

Is there any difference in the rates at which the gas is given off?

To what factors would you attribute the difference in the rates of reaction?

What must be the source of the gas in each case? Why?

2. Repeat Procedure 1 using (a) dilute hydrochloric acid, (b) dilute nitric acid.

Is hydrogen given off in each case?

Explain the action with nitric acid? (TEXT p. 361)

3. Into separate test-tubes, place (a) a small piece of pure zinc, (b) a small piece of impure (mossy) zinc. To each test-tube add a few millilitres of dilute sulphuric acid.

Describe the action that takes place in each test-tube.

To what could the difference be attributed?

4. To the test-tube containing the pure zinc and dilute sulphuric acid (Procedure 3), add a small crystal of blue vitriol (cupric sulphate hydrate).

What effect has the blue vitriol crystal on this reaction?

Name and account for the small brown deposit remaining in the tube after the reaction.

Can cupric sulphate be considered a catalyst in this reaction? Explain your answer.

5. To a test-tube containing a small piece of pure zinc and a small crystal of blue vitriol, add 1 ml. of concentrated sulphuric acid. Observe closely.

CAUTION: When water and concentrated sulphuric acid are mixed, so much heat is liberated that if the volume of water is small relative to the volume of concentrated acid, the heat converts the water into steam causing an explosion. The only safe way to dilute the acid is to add slowly the concentrated acid to the water. However, in this particular experiment, the following procedure may be followed if sufficient caution is exercised.

With a pipette or an eye dropper, allow one drop of water to fall into the tube, and after all action ceases, add one drop more. Repeat the addition of water, one drop at a time, and after each addition observe closely the reaction in the test-tube.

Is any hydrogen produced with pure zinc and concentrated sulphuric acid?

What happens when the first drop of water is added to the concentrated acid?

What is the effect of the gradual addition of water on the rate at which hydrogen is produced? Explain.

Explain why it is extremely dangerous to add more than one drop of water at a time to concentrated sulphuric acid.

6. Permit the action in Procedure 5 to continue for a short time, then filter the mixture and collect a small quantity of the filtrate (a) in an evaporating dish, (b) in a test-tube. Heat the evaporating dish strongly at first, but when the volume of liquid is considerably reduced, evaporate to dryness with a much smaller flame (*Technique 5*).

CAUTION: As the solution becomes more concentrated, small pockets of steam cause the liquid to spatter.

Put the test-tube containing the filtrate in your locker for a few days.

Why is it important to control the heating in this evaporation?

What is the colour of the residue? What is the chemical name for this substance? Account for its formation.

Describe the appearance of the contents of the test-tube after standing for a few days. Explain.

QUESTIONS

1. From information gained from these experiments, suggest the most easily controlled method of obtaining large volumes of hydrogen from dilute acids.

2. Write word equations representing the action of calcium, magnesium, zinc, and iron on the dilute acids used.

3. Do the results of Experiment 25 confirm the position in the activity series of the metals used?

4. Suggest a reason for not using sodium and potassium with dilute acids in the production of hydrogen.

EXPERIMENT 26. To prepare and collect hydrogen and study its properties.

PREPARATION AND COLLECTION

CAUTION: Do not strike a match or create a flame of any kind until so instructed.

1. Set up the apparatus as shown in Figure 34, and place a handful of mossy zinc in the flask. Add dilute sulphuric acid through the thistle-tube until the zinc is well covered, making sure that the lower end of the thistle-tube is covered by the acid.

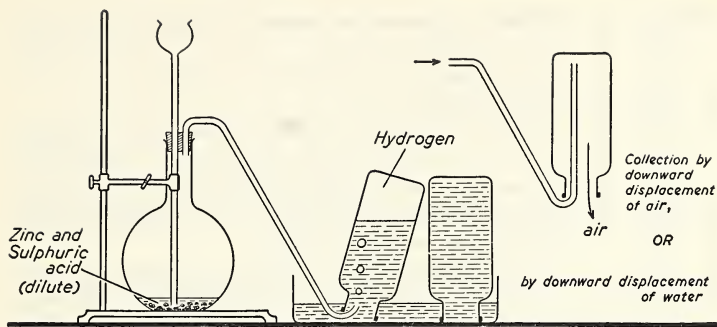


Fig. 34. The laboratory preparation and collection of hydrogen.

2. Let the reaction proceed for a short time and then collect, by the downward displacement of water, four test-tubes and one gas bottle full of the gas. Collect one full bottle of the gas by the downward displacement of air.

3. Stand the test-tubes and gas bottles, mouth downward, on the table until ready to use. Fill the generator with water to stop the production of hydrogen.

Why is it important to have the lower end of the thistle-tube covered with acid?

What two functions are served by the thistle-tube?

What precaution must be taken if the delivery tube is made of rubber tubing or contains rubber connections?

PROPERTIES OF HYDROGEN

1. To test the solubility of hydrogen in water, place one of the test-tubes full of gas, mouth downward, in a beaker of water, and allow it to stand until the end of the period. In determining the solubility, lift the test-tube so that the level of any water inside the tube is the same as that in the beaker.

Why is it necessary to adjust the water levels in determining the solubility of a gas in this way?

What does this experiment show about the solubility of hydrogen in water?

2. Look at the hydrogen generator to make sure that it has ceased producing hydrogen. You may now light a burner. Bring a blazing splint to the mouth of one of the remaining test-tubes of gas held in an inverted position, and carefully observe the manner in which the gas burns.

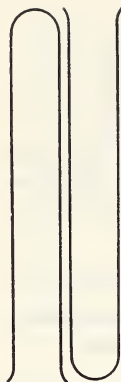


Fig. 35. Determining whether hydrogen is more or less dense than air.

What is characteristic about the way in which pure hydrogen burns?

Can hydrogen be identified by this characteristic behaviour in burning?

3. Lift the two remaining test-tubes from the table, quickly invert one of them, and hold the two tubes in the position shown in Figure 35 for at least half a minute. At the end of this time, bring a blazing splint to the mouth of each test-tube.

Describe the action in the test-tube held mouth upward.

Describe the action in the test-tube held with the mouth down.

What does this experiment indicate regarding the relative densities of hydrogen and air?

4. Lift the bottle of hydrogen, collected by the downward displacement of water, from the table, and keeping the mouth down (Fig. 36), slowly insert a long blazing wooden splint. Leave the splint in the bottle for a short time, and then slowly withdraw it.

Describe the action that takes place at the mouth of the bottle.

What happens to the burning splint? Does hydrogen support the combustion of the splint?

Name three important properties of hydrogen determined by this experiment.

5. Note the appearance of the interior of the remaining bottle of hydrogen collected by the downward displacement of air. Place the bottle, mouth upward, on the table, and immediately bring a blazing splint to the mouth. Observe the nature of the flame and any deposit on the inside of the bottle after the hydrogen has burned.

What is the reason for placing the mouth of the bottle upward in this experiment?

Describe the flame and the appearance of the interior of the bottle after the gas has burned.

Suggest a name for the substance appearing inside the bottle.

Suggest a method of obtaining a sufficient amount of this substance to make a positive identification.

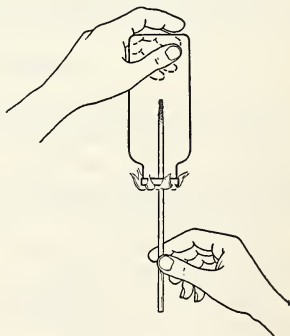


Fig. 36. Thrusting a blazing splint into a bottle of hydrogen.

Why was this particular bottle collected by the downward displacement of air?

QUESTIONS

1. Write a word equation for the laboratory preparation of hydrogen.
2. List five properties of hydrogen.

EXPERIMENT 27. (*Demonstration*) The burning of dry hydrogen at a jet.

1. Set up the apparatus as shown in Figure 37, making sure that there is a plentiful supply of zinc in the flask and an additional supply of acid available in case the supply of hydrogen dwindles. **NOTE:** The hydrogen generator shown in the diagram may be replaced by a Kipp apparatus (TEXT p. 62), which will supply a steady stream of hydrogen for a longer time than will the flask-type generator.

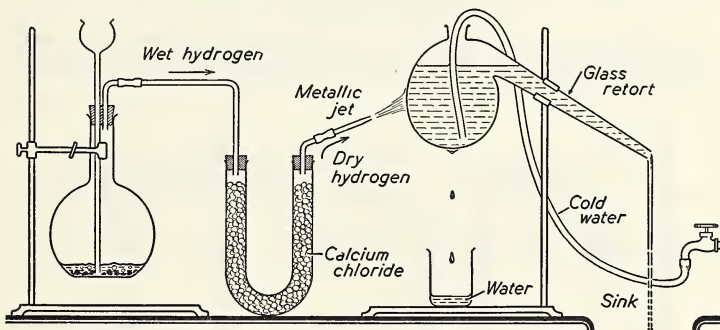


Fig. 37. Burning dry hydrogen at a jet.

2. Let the reaction progress for a short time so that the escaping hydrogen will displace the air already in the system. To be certain that pure hydrogen is issuing from the jet, collect a test-tube full of the gas, and keeping the mouth down, bring it to the flame of a burner placed at some distance from the generating apparatus. When the gas burns quietly, with no explosion, bring it to the jet and ignite the escaping hydrogen with this flame. **This is the only safe method of lighting a hydrogen jet.**

Explain the importance of this method of lighting the jet.

3. Let the flame of the burning hydrogen contact the surface of the glass retort, and observe the surface closely. Test the substance formed with anhydrous cupric sulphate.

Describe the flame of the burning hydrogen.

Describe the appearance of the outer surface of the retort.

What does the anhydrous cupric sulphate test indicate?

Account for the formation of the deposit on the flask.

QUESTIONS

1. Why is it important to dry the hydrogen in Experiment 27? Does the calcium chloride in the tube show any evidence of being effective as a drying agent?

2. How does this experiment illustrate the *qualitative synthesis* of water?

3. Write a word equation to represent the burning of hydrogen in air.

EXPERIMENT 28. (*Demonstration*) The action of dry hydrogen as a reducing agent.

1. Arrange the apparatus as in Figure 38. Spread the cupric oxide evenly along the tube, and make certain that the glass delivery

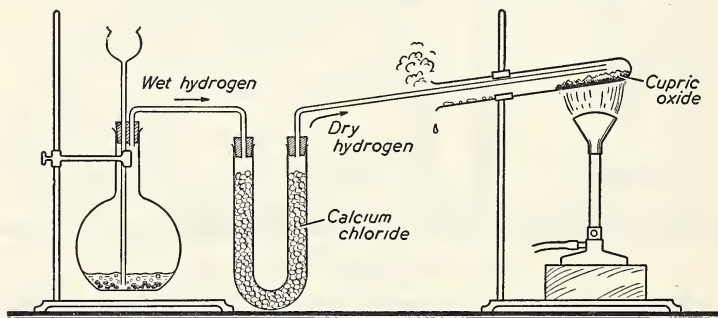


Fig. 38. Reducing cupric oxide with dry hydrogen

tube does not become clogged when placing it in the test-tube. Add dilute sulphuric acid to the zinc in the flask, and let the reaction continue for a sufficient length of time to drive the air out of the apparatus. Then bring the flame under the part of the tube containing the cupric oxide. Do not bring the burner flame near the open end of the test-tube. Continue heating until the action within the test-tube ceases. Then remove the burner, and carefully withdraw the glass tube, making sure that it does not touch the sides of the test-tube.

Give two reasons for keeping the burner flame away from the mouth of the test-tube.

What changes take place in the cupric oxide? (TEXT p. 65)

What is the nature of the deposit near the mouth of the test-tube?

List any other observations noted.

2. Allow several drops of the liquid formed in the test-tube to fall on some anhydrous cupric sulphate. When the test-tube has cooled sufficiently, tap the material in the tube onto a watch-glass, and after observing the material closely, add a few drops of concentrated nitric acid. Rinse out the test-tube with water, examine any deposit on the inside of the tube, and add a few drops of concentrated nitric acid.

What is the liquid formed in the cool part of the test-tube? Account for its formation.

What is the appearance of the mixture on the watch-glass?

What changes take place when the concentrated nitric acid is added?

What does this test indicate? (TEXT p. 361)

Name three possible substances present in the mixture.

What is the deposit adhering to the inside of the test-tube? Account for its formation.

QUESTIONS

1. Write two possible word equations for the action of hydrogen on cupric oxide.

2. Explain the action of hydrogen as a reducing agent in this experiment.

UNIT 5

Water

EXPERIMENT 29. The purification of water.

1. Add a tablespoonful of garden soil to 500 ml. of water in a beaker, and stir. Pour half of the mixture into another beaker so that you now have two beakers containing muddy water. Allow one beaker to stand for 20 minutes.

2. Add 10 ml. of a 10% solution of aluminum sulphate to the second beaker, and stir. Then add 10 ml. of a saturated solution of calcium hydroxide (limewater), stir, and allow to stand.

Compare the rates of settling (sedimentation) of the soil particles in Procedures 1 and 2. Explain the difference in rates.

3. Stir the water in the first beaker again, and filter a small quantity of it. Compare the clarity of the filtrate with that of the water above the sediment in the untreated beaker.

Which is clearer: filtered water, water from which the sediment has settled by gravity, or water which has been clarified by chemical action?

QUESTIONS

1. What effect has the addition of certain chemicals on the speed of settling or sedimentation? Explain how this effect is produced.

2. Of the above three methods, which is the most efficient from the standpoint of the clarity of the water produced? Which is most efficient from the standpoint of the time required?

3. Is the water, produced either by sedimentation or by filtration, actually pure water? Explain.

4. Would any of the above methods make water safe for drinking? Explain.

EXPERIMENT 30. (*Demonstration*) The distillation of water.

1. Into a beaker containing 50 ml. of water, put a few crystals of potassium permanganate, and stir. Add a teaspoonful of soil, and stir again.

What happens to the potassium permanganate? Was the water in the beaker clear after the potassium permanganate was added?

Was the water clear after the soil was added?

2. Set up a distilling apparatus as shown in Figure 39. Pour the mixture from Procedure 1 into the flask (A), connect the lower rubber tubing attached to the condenser to a water tap, and lead the other

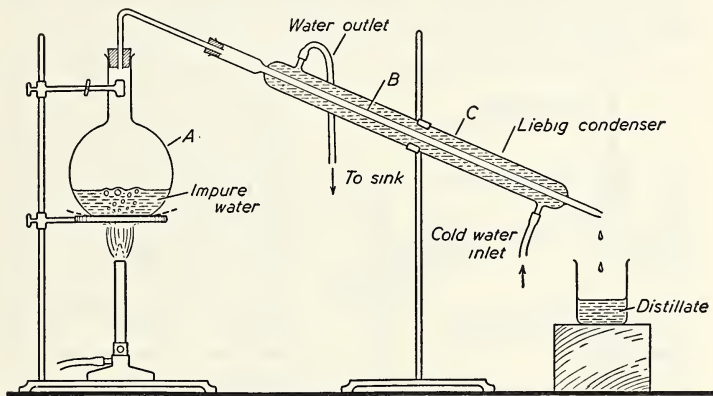


Fig. 39. The distillation of water using a Liebig condenser.

tubing to the sink. Turn on the tap so that a steady stream of cold water flows through the glass jacket (C) of the condenser. Heat the mixture in the flask until it boils, and continue to heat gently for ten minutes. Collect some of the distillate in a beaker as shown. Examine and taste the distillate. Evaporate about 2 ml. of the distillate to dryness in either an evaporating dish or a Pyrex test-tube.

Compare the properties of the mixture in the flask with the properties of the distillate.

What is observed when some of the distillate is evaporated to dryness? What does this show about the effect of distillation on (a) dissolved, and (b) suspended substances?

Name the two changes of state involved in the distillation process. Is the distillate pure water? Explain.

Why is the cold tap water made to enter at the lower end of the condenser?

QUESTIONS

1. Compare water obtained by the process of distillation with water obtained by filtration as to (a) clarity, (b) purity.

2. Why is distillation often called a *double* process?
3. Why is the process of distillation not used in obtaining a pure water supply for a city?

EXPERIMENT 31. (*Demonstration*) Fractional distillation.

Replace the flask used in Experiment 30 with a distilling flask which has a side-arm, and fit it with a thermometer as shown in

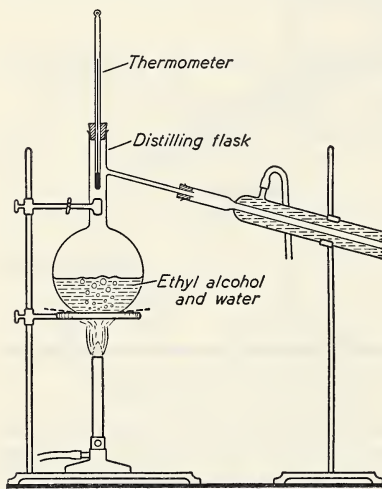


Fig. 40. The fractional distillation of a mixture of ethyl alcohol and water.

Figure 40. Pour a mixture of equal volumes of ethyl alcohol and water into the flask, and then connect the condenser as shown. Heat the flask until the liquid boils, and continue to heat slowly for ten minutes. Take temperature readings every two minutes. Collect five separate samples of the distillate, and smell each sample.

At what temperature does the initial boiling occur? Is there any change in the thermometer reading as the boiling continues?

Which of the two liquids in the distilling flask boiled at

the initial temperature?

Which distillate contains the highest percentage of ethyl alcohol (TEXT p. 76)

What name is given to the process whereby two liquids of different boiling points are separated?

How could you prove that each sample of distillate differs considerably from the liquid originally in the distilling flask?

QUESTIONS

1. Explain the use of the word *fractional* in referring to a distillation of this type.
2. Explain briefly how a mixture of liquids, such as are found in crude oil, may be separated by fractional distillation. (TEXT p. 283)

EXPERIMENT 32. The properties of water.

1. Recall experiments performed in earlier grades, and record each of the following properties of water in your notebook: colour, odour, taste, conductivity of heat and of electricity, freezing point (760 mm. pressure), boiling point (760 mm. pressure), heat of vaporization, heat of fusion, and density at 4°C.

2. Review your observations for the experiments completed in Unit 2 (Oxygen) and Unit 4 (Hydrogen). In this review, concentrate on the properties of water, and list all the chemical properties illustrated by these experiments.

3. Add a pinch of anhydrous cupric sulphate to a few drops of water on a watch-glass.

Describe the appearance of anhydrous cupric sulphate, (a) before water is added, (b) after water is added.

When water is added to anhydrous cupric sulphate, is there a physical or chemical change? Explain.

EXPERIMENT 33. A study of hydrates.

1. Gently heat a few crystals of cupric sulphate hydrate (bluestone) in a clean, dry test-tube clamped to a retort stand, with the mouth slightly lower than the rest of the tube. Clamp on a second test-tube vertically, as shown in Figure 41. Increase the temperature gradually until the crystals are being strongly heated. Test a drop of the material caught in the vertical tube with anhydrous cupric sulphate.

What two changes are observed in the crystals of cupric sulphate hydrate?

Describe the substance found in the vertical test-tube and on the cooler portion of the sides of the heated test-tube. Account for its formation. What is this liquid? How did you identify it?

Does the heating of a hydrate represent a physical or a chemical change? Explain.

2. Gradually move the Bunsen burner forward towards the mouth of the upper test-tube, and continue heating until the tube and its contents are dry. Cool the test-tube. Shake out the contents onto

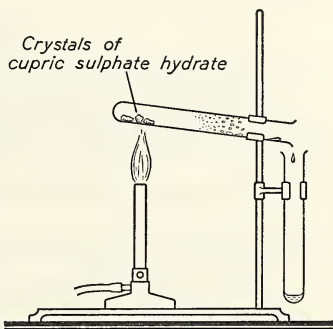


Fig. 41. The dehydration of bluestone.

a piece of paper, and examine the substance. Add a drop of water to this substance.

Describe the appearance of the substance on the paper.

What is observed when water is added to the substance formed by heating bluestone? What is this substance?

Write a word equation to represent the reaction that takes place when bluestone is strongly heated. What name is given to this type of reaction? (TEXT p. 78)

Write a word equation to represent the reaction that takes place when water is added to anhydrous cupric sulphate. What name is given to this type of reaction?

3. Repeat Procedures 1 and 2 using a few crystals of each of the following substances in separate dry test-tubes: alum, sodium chloride, and sodium carbonate.

Describe what happens in each case.

Which of these substances decomposes when heated, to form an anhydrous salt and water of hydration? What are such substances called?

Would this experiment indicate that all crystalline substances are hydrates?

QUESTIONS

1. Define a hydrate, and name five examples.
2. Why is *water of hydration* more suitable to express the water present in a hydrate than is *water of crystallization*?
3. Suggest an experiment that could be performed to find the percentage of water of hydration in bluestone.

EXPERIMENT 34. Efflorescence and deliquescence.

1. Weigh about a teaspoonful of sodium thiosulphate crystals (*hypo*) on a watch-glass, and allow them to stand for most of the laboratory period. Then weigh again.

Describe any change in the appearance of the crystals.

Account for any change in weight observed.

2. Weigh about a teaspoonful of granular anhydrous calcium chloride on a watch-glass, and allow it to stand for most of the laboratory period. Then weigh again. (While waiting, continue with Procedure 3.)

Does the calcium chloride change in appearance? If so account for the change.

Does the weight of the calcium chloride change? If so why?

3. (For this procedure, the class may be divided into groups.) On

separate watch-glasses, place about one-quarter of a teaspoonful of each of the following nine substances; "hypo" crystals, granules of anhydrous calcium chloride, pellets or lumps of sodium hydroxide, crystals of sodium nitrate, crystals of potassium nitrate, crystals of magnesium chloride, washing soda crystals, crystals of Glauber's salt, and anhydrous cupric sulphate. Place the watch-glasses and their contents in a convenient place until the next laboratory period, and examine the substances again.

Describe any changes that occur in each of the substances used.

Compare the "hypo" after exposure to air for a day with the "hypo" exposed for 20 minutes, as in Procedure 1.

What other substances behave as did "hypo"? What are such substances called? (TEXT p. 78)

Compare the calcium chloride after exposure to air for a day with the calcium chloride exposed for 20 minutes, as in Procedure 2.

What other substances behave as did calcium chloride? What are such substances called? (TEXT p. 78)

How do you explain the behaviour of the anhydrous cupric sulphate?

QUESTIONS

1. Distinguish between *efflorescence* and *deliquescence*. Are both of these phenomena types of chemical reactions? Explain.
2. Explain why deliquescent substances make good drying agents.
3. Distinguish between *dehydration* and *efflorescence*.

EXPERIMENT 35. (*Demonstration*) The electrolysis of water.

1. A Hoffman apparatus is used as shown in Figure 42. To about 250 ml. of water add about 5 ml. of concentrated sulphuric acid (*Technique 3*). With both stop-cocks open, slowly pour the acidulated water into the reservoir until the two outer tubes are filled up to the taps. Avoid getting any liquid in the jets above the taps. Close the taps, and add enough acidulated water to half fill the reservoir. Paste a paper label on the reservoir so that its upper edge marks the surface level of the liquid it contains. Connect the platinum electrodes to a source of *direct* current. If 3 dry cells in series are used, the electrolysis should be started before the class period to ensure the accumulation of enough gases for testing. If a D.C. source of higher voltage (e.g., 110 volts) is used, a lamp or some other resistance must be connected in series in the circuit. The circuit, which includes a switch and the Hoffman apparatus, should be carefully examined and the direction of the current determined, before beginning the experiment.

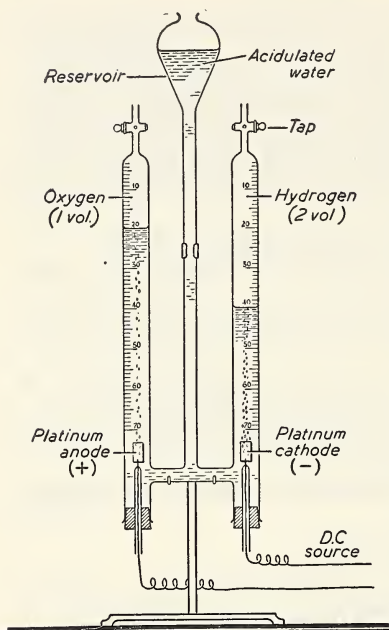


Fig. 42. The electrolysis of water using a Hoffman apparatus.

Close the switch and allow the current to pass through the acidulated water until the tube above the cathode contains about 20 ml. of gas. Open the switch, and accurately read the volume of gas collected (a) above the cathode, and (b) above the anode. Observe the new level of the liquid in the reservoir. Hold an inverted test-tube over the jet above the cathode, and open the tap. Keep the tap open until all the gas escapes into the test-tube, but do not permit any liquid to escape from the jet. Bring a lighted splint to the mouth of the inverted test-tube.

Test the gas above the anode by holding a glowing splint above the jet, at the same time opening the tap. Do not allow any liquid to escape from the

jet. Observe the level of the liquid in the reservoir after the tubes are again filled with liquid.

Describe what was observed when the switch was closed.

What was the volume of the gas collected (a) above the cathode, (b) above the anode?

Name the gas (a) above the cathode, (b) above the anode. How were these gases identified? What were the relative volumes of the two gases?

From a comparison of the liquid levels in the reservoir before the electrolysis began, and after the tubes were again filled with liquid, what conclusion may be made about the amount of water that has been electrolysed?

2. If time permits, repeat the above experiment.

QUESTIONS

1. Make a diagram of a Hoffman apparatus connected in a circuit, as used for the electrolysis of water. Label all the essential features, and include a word equation for the action that takes place.

2. In this experiment, why (a) was *direct current* used, (b) was a small quantity of sulphuric acid added, (c) were platinum electrodes used?

3. Explain the changes in level of the acidulated water in the reservoir, as the electrolysis took place, and as the gases were allowed to escape from the jets.

4. What would have been observed if an *alternating current* had been used in this experiment?

5. Why is the electrolysis of water an *analysis by volume*?

EXPERIMENT 36. (*Demonstration*) The synthesis of water (by volume).

1. A eudiometer tube is used (Fig. 43). Because of the expense of the large volume of mercury required to fill the jar and tube, and the greater danger of breakage, it is suggested that water be used in place of mercury. Examine a eudiometer tube so that you will be able to describe it in detail. Fill the tube with water and invert it in a tall hydrometer jar also filled with water. Pass a small volume (e.g., 15 ml.) of hydrogen (from a hydrogen cylinder or generator) into the tube.

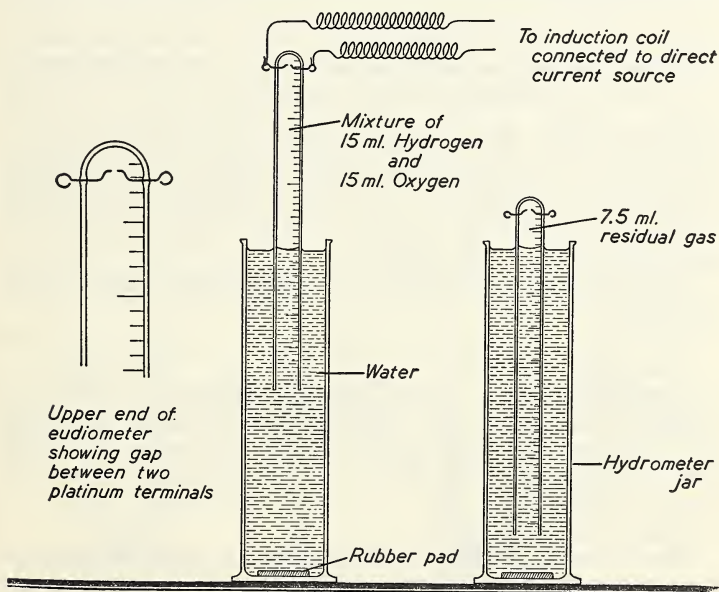


Fig. 43. Synthesis of water by volume in a eudiometer.

Lower the tube until the level of the water is the same inside and outside. Read the volume of hydrogen accurately, and record it. Add about the same volume (15 ml.) of oxygen (from an oxygen cylinder or generator), and again adjust the tube until the water levels are the same inside and outside. Read and record the volume of the mixed gases.

CAUTION: This is a dangerous experiment. Because the eudiometer may shatter, no one except the instructor should be within 10 feet of the demonstration table. The eudiometer may be made "shatterproof" by cementing several layers of cellophane on the glass without affecting visibility. Further protection may be had by using a fine-mesh wire screen as a shield for both pupils and teacher.

It is imperative that the two gases be present in volumes that assure some residual gas in the eudiometer after the explosion. The mixture of 15 ml. of hydrogen and 15 ml. of oxygen represents the maximum volume that should be used in a 50 millilitre eudiometer. Possibly 10 ml. of hydrogen and 10 ml. of oxygen would be a safer choice of volumes for the average experiment.

Connect the platinum terminals to an electric circuit containing three or four dry cells connected in series, an induction coil, and a switch (open). Grasp the eudiometer with a towel and press it firmly on a rubber pad placed in the bottom of the hydrometer jar, and at the same time support the hand on the top of the jar. Close the switch. Observe any changes that occur inside the tube at the instant the switch is closed and immediately afterwards. Wait until the tube cools. Then adjust the levels as before, and measure the volume of the residual gas. Pass a spark through the residual gas.

What is observed at the moment the circuit is closed? Explain the downward movement of the water.

What caused the sudden uprush of water that followed shortly after the explosion?

Why does no explosion occur when a spark is passed through the residual gas?

2. Place the hand or a small glass plate under the mouth of the eudiometer tube, which now contains only the residual gas and water, and remove the tube and contents from the hydrometer jar. Place the tube in an upright position; remove the hand and immediately insert a glowing splint.

What is the residual gas?

What is the ratio by volume of the hydrogen and oxygen which disappeared? What happened to these gases? What was observed during the experiment to support your conclusion?

3. Repeat the experiment, but this time use approximately 20 ml. of hydrogen and 7 ml. of oxygen. Tabulate your data as follows:

1. Volume of hydrogen used.....ml.
2. Volume of the mixture of hydrogen and oxygen.....ml.
3. Volume of oxygen used (2-1).....ml.
4. Volume of residual gas.....ml.

4. After the explosion has taken place and the volume of the residual gas has been measured accurately, add about 5 ml. of hydrogen to the eudiometer tube. Pass a spark through the gas.

Explain why no explosion occurs.

Suggest how you might verify your identification of the residual gas.

QUESTIONS

1. What would be one advantage of using mercury to fill the tube and jar, in preference to water?

2. Was the reaction between oxygen and hydrogen endothermic or exothermic? Give reasons for your answer.

3. Why should care be taken not to pass hydrogen and oxygen into the eudiometer tube in the exact proportion in which they unite? (TEXT p. 84)

EXPERIMENT 37. (*Demonstration*) The synthesis of water (by weight). (*Dumas' experiment*)

The apparatus used (Fig. 44) consists of three drying tubes containing granular calcium chloride, a combustion tube and porcelain boat, a Bunsen burner fitted with a flame spreader, and a hydrogen generator (not shown in the diagram). Weigh the drying tubes (C) and (D) together, and record their weight. Fill the boat almost full of cupric oxide, place it in the combustion tube, and weigh the tube and contents together. Record this weight. Assemble the

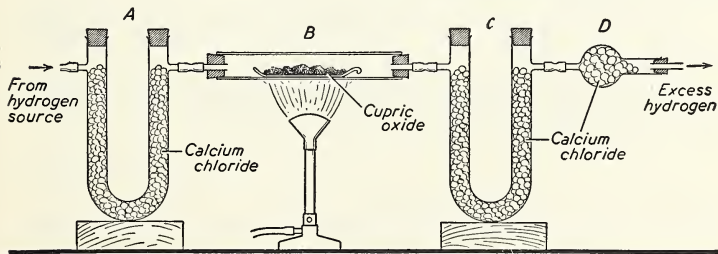


Fig. 44. Finding the percentage composition of water by weight.

complete apparatus, and connect it to a hydrogen generator. Pass hydrogen through the assembled tubes until all the air has been driven from them. To make certain that hydrogen has replaced the air in the system, test at the exit tube, for pure hydrogen (see Experiment 27). When hydrogen only is escaping, then, **and not until then**, heat the combustion tube as shown in Figure 44. Heat strongly for 5 minutes. Then remove the flame, and allow the tube and contents to cool. Remove (*C*) and (*D*) together, and weigh them again. Remove the combustion tube with its contents, and weigh again. Record the observations as follows:

1. Weight of (*B*) before reduction.....gm.
2. Weight of (*B*) after reduction.....gm.
3. Weight of tubes (*C* + *D*) before experiment.....gm.
4. Weight of tubes (*C* + *D*) after experiment.....gm.

From these observations *calculate*:

- (a) Weight of oxygen used ($1 - 2$).....gm.
- (b) Weight of water formed ($4 - 3$).....gm.
- (c) Weight of hydrogen in the water formed ($b - a$).....gm.

Explain why ($1 - 2$) gives the weight of the oxygen used in forming the water obtained.

Explain why ($4 - 3$) gives the weight of the water that is formed by the reduction.

From the available data calculate the percentage composition of water by weight.

QUESTIONS

1. Why was the drying tube (*A*) used? Why was it not weighed before and after the reduction?

2. Why was the heating of the combustion tube delayed until the tubes were free of air?

3. What is the percentage error in your experiment? Compare your result with that of Morley. (TEXT p. 86)

UNIT 6

Solutions and Mechanical Mixtures

EXPERIMENT 38. The formation of solutions.

1. Add a small crystal of potassium permanganate to 10 ml. of water in a test-tube, and shake until the crystal disappears. Note the appearance of the liquid. Add a few more small crystals, and shake again until the crystals disappear.

What is the colour of a solution of potassium permanganate and water?

Is the solution clear or turbid?

What effect has the addition of more crystals on the colour of the solution?

Is there any effect on the clarity of the solution?

2. Place 50 ml. of water in a graduated cylinder. On a piece of paper weigh 6 gm. of ammonium chloride. Take and record the temperature of the ammonium chloride and then of the water. Add the ammonium chloride to the water, and shake until solution is effected, making sure that none of the solution escapes. Take the temperature of the solution and read its volume.

What is the colour of the ammonium chloride solution?

Is there a change in temperature? If so, state the change.

Is there any change in the volume of the liquid?

3. (*Demonstration*) Place about 75 ml. of water in a gas-measuring tube (capacity 200 ml.). Slant the tube at an angle, and slowly add about 75 ml. of alcohol in such a way that very little mixing of the two liquids occurs. Hold the tube upright, and record the volume accurately. Put a stopper in the tube, and shake for a few minutes. Record the volume.

What evidence is there that a solution has been formed?

What was the total volume of the liquids before mixing? What was the volume after mixing?

Does the solution occupy as great a volume as the two components occupy separately?

4. Allow water to run from the cold-water tap until it is at the coldest possible temperature. Fill a beaker with cold water, and set it aside in a warm place for several hours. Examine the inside walls of the beaker.

Account for the bubbles that are seen clinging to the walls of the beaker. What gases are present in these bubbles?

What does this experiment indicate about the relative solubilities of air in cold water and in warmer water? Is this property characteristic of all gases? (TEXT p. 95)

EXPERIMENT 39. Factors affecting the rate of solution.

1. Half fill two test-tubes with water. To each, add 1 gm. of cupric sulphate crystals, and allow one tube to stand undisturbed. Cover the mouth of the second tube with your thumb and shake vigorously.

Is cupric sulphate soluble in water? Give reasons for your conclusion.

What effect has agitation on the rate at which the distribution of the solute throughout the solvent takes place? Explain.

Do you think that the cupric sulphate in the undisturbed test-tube would eventually dissolve? Explain.

2. Half fill two test-tubes with water. To one, add 2 gm. of cupric sulphate crystals, and to the other the same quantity of powdered cupric sulphate made by grinding the crystals with a mortar and pestle. Shake each tube successively, and note the time required to dissolve the cupric sulphate in each.

Which dissolves more rapidly? Why is there a difference in the rate of solution?

What method of speeding up the rate of solution is illustrated by this experiment?

3. Half fill two test-tubes with water. Add 2 gm. of cupric sulphate crystals to each. Let one stand undisturbed, and warm the other in a Bunsen flame.

What effect has an increase in temperature on the rate of solution?

EXPERIMENT 40. Factors affecting solubility.

1. Half fill six test-tubes with water, and place them in a large beaker or a test-tube rack. To each test-tube, add one of the following substances: 1 gm. of cane sugar, 1 gm. of sodium chloride, 1 gm. of calcium carbonate, 2 ml. of denatured alcohol, 2 ml. of glycerine, and 2 ml. of kerosene. Shake each tube, and observe the results.

Which of these substances are insoluble? Which are soluble in water? Which are immiscible? Which are miscible?

Which of the following factors is responsible for the difference in solubilities: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

2. Half fill three test-tubes with water, alcohol, and carbon tetrachloride respectively. Add a small crystal of iodine to each, shake, and observe the results.

Does the iodine dissolve in all three liquids? Explain.

Which of the following factors is responsible for the difference in solubility: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

3. Repeat Procedure 2 using pea-sized lumps of margarine in place of iodine crystals.

Does the margarine dissolve in all three liquids? Explain.

Which factor is responsible for the difference in solubility: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

Which of the three liquids used would be most satisfactory for removing grease stains from cloth? Why?

4. Weigh 10 gm. of ammonium chloride on a piece of paper, and then divide it into five portions of approximately 2 gm. each. Put 10 ml. of water into a large test-tube, add a 2 gm. portion of the salt, and shake the test-tube until the salt dissolves. Repeat this process until a small amount of the solute remains undissolved. Heat the test-tube until the solution is warm (about $50^{\circ}\text{C}.$), and note the effect on the excess salt. Add another portion of the salt and shake again. Heat again until the solution boils. Add the remaining portions of the salt, shaking after each addition. Cool to room temperature.

Which factor is responsible for the difference in solubility: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

How does temperature affect the solubility of ammonium chloride?

Which of the solutions examined were unsaturated? Which were saturated? Explain.

What occurred when the hot concentrated solution was cooled? Explain.

QUESTIONS

1. Distinguish between *rate of solution* and *solubility*.

2. A salt is shaken in water for some time, and when agitation ceases, some of the salt settles. What is the condition of the solution above the salt? How may the excess salt be put into solution?

3. What four factors are always implied when the term *solubility* is used? Define *solubility*. (TEXT p. 96)

4. Suggest how you would determine the solubility of ammonium chloride at $60^{\circ}\text{C}.$

5. Distinguish between the terms *clear* and *colourless*. Are all solutions clear? Are all solutions colourless?

6. What is the chief distinguishing characteristic of all solutions?

EXPERIMENT 41. The determination of the solubility of a salt.

It is suggested for this experiment that the solubility of several salts be determined by distributing various salts to different groups. The method given here for common salt (sodium chloride) would apply, except for quantities stated, to each of the salts selected.

1. Shake together in a *large* test-tube 30 ml. of water and 15 gm. of sodium chloride until no more salt dissolves. Take the temperature of the saturated solution.

2. Weigh together a clean, dry evaporating dish and a watch-glass cover. Record this weight. Filter about 20 ml. of the saturated solution into the evaporating dish, place the cover on it, and weigh. Record this weight. Remove the cover, and place the evaporating dish over a beaker of boiling water (or on a sandbath)

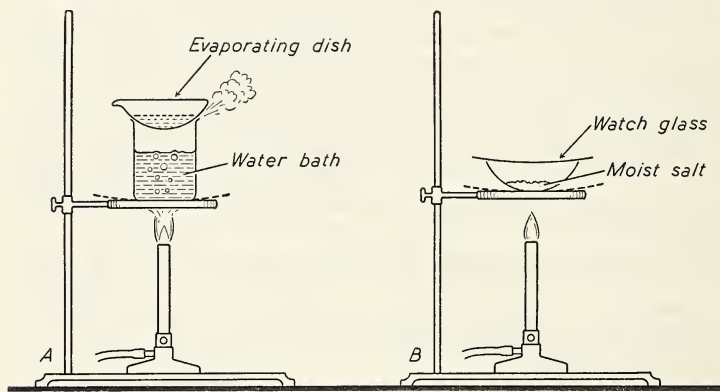


Fig. 45. Evaporating a solution. *A*, first stage, using a water bath. *B*, second stage, evaporating to dryness.

to evaporate the solution (Fig. 45*A*). Evaporate until nearly dry, and then remove the evaporating dish. Wipe the water from the outside of the dish, and then place it on a wire gauze supported on a ring stand.

3. Again cover the dish with the watch-glass, and heat it carefully with a small Bunsen flame (Fig. 45*B*) until thoroughly dry. Be sure that all traces of moisture are driven not only from the evaporating dish, but from the under surface of the cover as well. Allow the dish to cool to room temperature, and then weigh the dish, contents, and cover together, and record the weight. Record your data as follows:

1. Temperature of the saturated solution °C.
2. Weight of evaporating dish and cover gm.
3. Weight of evaporating dish and cover and solution gm.
4. Weight of saturated solution (3—2) gm.
5. Weight of evaporating dish and cover and residue gm.
6. Weight of sodium chloride dissolved (5—2) gm.
7. Weight of water in the saturated solution (4—6) gm.

Calculate the weight of sodium chloride required to saturate 100 gm. of water at the temperature of the solution in your experiment.

How does your result compare with that shown by the solubility curve on page 97 of the text?

What is your percentage error?

Define the solubility of a salt.

QUESTIONS

1. Why was the evaporating dish covered with a watch-glass during the last stage of evaporation?

2. What sources of error may have caused your result to be inaccurate?

3. Suggest how you could determine the solubilities of a certain salt at different temperatures.

EXPERIMENT 42. The preparation of a supersaturated solution.

1. Put crystals of sodium thiosulphate (hypo) to a depth of one inch in a clean, dry test-tube. Add 10 (no more) drops of water. Place a thumb over the end of the tube and shake it for a minute. Note any change in temperature by placing the test-tube against the back of your hand. Examine the contents.

What evidence is there that "hypo" is soluble in cold water?

Is there any change in temperature when solution occurs?

2. Put crystals of sodium thiosulphate (hypo) to a depth of four inches in a clean, dry test-tube, taking care that no crystals stick to the sides of the tube (**Technique 8**). Add 10 (no more) drops of water, and gently warm the tube (without agitation) until all the crystals are in solution. Place some absorbent cotton in the mouth of the test-tube, and hold the tube in a stream of cold water from the tap until it is cooled to approximately its initial temperature. Examine the contents of the tube.

At what stage during this procedure is the solution saturated? Give a reason for your answer.

Is the hot solution saturated before cooling? How could you have tested it to find out?

Why is the cold solution you have just made called a supersaturated solution?

3. Hold the test-tube that contains the supersaturated solution so that it may be viewed against a light background. Add first a crystal of blue vitriol, and one minute later, a small crystal of "hypo", and observe. Note any temperature change.

What was the effect of adding the cupric sulphate hydrate crystal?

What happened when the "hypo" crystal was added? Was there any change in temperature? What kind of solution remained?

From knowledge gained in previous experiments, why would you expect to have an increase in temperature when crystallization occurs?

QUESTIONS

1. In preparing the supersaturated solution of "hypo", why was absorbent cotton used as a plug in the end of the test-tube?

2. Account for the presence of such a large volume of solution when only a few drops of water were used in the preparation of the supersaturated solution.

3. Why was a crystal of a different solute added to the supersaturated solution of sodium thiosulphate?

4. When a crystal of the same solute is added to a supersaturated solution, what is the condition of the solution that results? Explain.

5. Describe, using diagrams to illustrate your answer, how you would make a supersaturated solution of Glauber's salt. What would happen if a small crystal of sodium sulphate were added to the solution?

EXPERIMENT 43. (*Demonstration*) The formation of a large crystal of alum.

1. Dissolve about 30 gm. of powdered potassium aluminum sulphate in 100 ml. of warm water (about 50°C.) in a beaker. Allow the solution to stand over night, and then filter it into another clean beaker. Select the most perfect crystal that has formed during the night, and place it in the clear filtrate. Cover the beaker with a filter paper, and again allow the solution to stand over night. On the following day remove the crystal, and clean it by scraping off any small nodules that may be adhering to it and then wiping it with a towel. Filter the solution to remove any new crystals that may have formed, and then replace the cleaned crystal in the filtered solution.

2. The above procedure should be repeated every day for about two weeks. The crystal should be turned from one crystal face to

another, at intervals, so that the growth of the crystal will be uniform.

After the final cleaning, if the crystal is to be handled, it may be covered with a coating of shellac.

EXPERIMENT 44. The distinction between a mechanical mixture and a chemical compound.

1. Review the properties of sulphur and iron as determined in Experiments 2 and 4. (TEXT p. 99)

List the physical properties of sulphur and those of iron.

2. Place a level teaspoonful of fine degreased iron filings and a level teaspoonful of powdered sulphur in separate piles on a piece of paper. Push the bulb of a thermometer into each pile in turn, and read the temperature of each substance.

Using a spatula, mix the iron and sulphur thoroughly. Take the temperature of the mixture, and examine it with a magnifying glass.

(Iron filings can be degreased by washing them several times in either carbon tetrachloride or alcohol, until the decanted liquid is almost colourless. They should then be washed with ether and allowed to dry by exposure to air.)

Did any temperature change occur when the two elements were mixed?

What is the colour of the mixture?

Can the iron and sulphur in the mixture be visibly distinguished?

3. Place about *one-fifth* of the mixture on a piece of paper, and draw a magnet through it.

Describe what happens. Can the iron and sulphur be separated by this method?

4. **CAUTION: Keep flames away. Carbon disulphide is volatile and very flammable.**

Put about *one-fifth* of the original mixture in a test-tube, and add about 2 ml. of carbon disulphide. Shake well and filter.

Catch the filtrate on a watch-glass, and allow it to evaporate. Examine the residue on the filter paper. Also examine the residue left on the watch-glass after the evaporation of the carbon disulphide.

Describe the residue on the filter paper. How could you test this residue so as to identify it positively?

Describe the residue on the watch-glass, and identify it.

What property of sulphur and what property of iron is shown by this experiment?

5. Drop a pinch of the mixture into a beaker of water, stir, and observe what happens.

State the reason for the difference in the behaviour of the sulphur and the iron.

6. Drop about one-third of the remaining mixture into a test-tube containing about 2 ml. of dilute hydrochloric acid. Apply a lighted match to the mouth of the test-tube.

Is there any evidence of a chemical reaction?

Is there any odour to the gas produced? Does it burn? Identify it.

What substance is left in the test-tube? Why?

7. CAUTION: (Demonstration) The reaction may be violent.

Place the remainder of the mixture in a test-tube, and heat gently until the mixture glows. Then immediately withdraw the flame.

Is there any evidence that a chemical change is taking place? Describe.

8. When the substance in the test-tube has cooled, break the tube in a mortar and remove the contents. Remove the broken glass, and crush the contents in the mortar with a pestle. Examine the powder with a lens.

What is the colour of the substance? Can you visibly distinguish the iron and sulphur?

9. Remove about one-fifth of the material from the mortar, and shake it with 2 ml. of carbon disulphide in a test-tube. Filter, collect the filtrate on a watch-glass, and allow it to evaporate.

Is there any residue? Is the substance sulphur?

10. Remove about half of the material remaining in the mortar, and draw a magnet through it.

Is the substance magnetic? Is it iron?

11. Place the remainder of the substance from the mortar into a test-tube, and add about 2 ml. of dilute hydrochloric acid.

Is there any evidence of a chemical reaction?

Describe the odour of the gas produced. Is it the same gas as that formed when the mixture of iron and sulphur was dropped into hydrochloric acid? Name the new solid compound formed. (TEXT p. 300) Name the gas produced. (TEXT p. 302)

QUESTIONS

1. Give four characteristics of the mechanical mixture of iron and sulphur, and four characteristics of the chemical compound ferrous sulphide.

2. Was the reaction that took place when the iron and sulphur combined endothermic or exothermic? State the evidence for your conclusion.

3. Consult the TEXT, p. 115, to find an additional difference (not shown by this experiment) between a mechanical mixture containing iron and sulphur and a chemical compound composed of iron and sulphur.

EXPERIMENT 45. The formation of a suspension.

1. Fill four test-tubes three-quarters full of water. Into each of two of these test-tubes, place as much precipitated chalk (calcium carbonate) as will lie on a Canadian five-cent piece. Into each of the two remaining test-tubes, place similar amounts of either powdered clay or finely divided soil.

Shake all four test-tubes vigorously, an equal number of times, and stand them in a test-tube rack for observation.

What is the appearance of the precipitated chalk mixture?

What is the appearance of the powdered clay mixture?

2. Let the test-tubes stand for ten minutes, and again observe the appearance of each.

Name two ways in which the appearance of the test-tubes containing clay differ from the appearance of the test-tubes containing precipitated chalk.

3. Filter separately the contents of (a) one of the test-tubes containing precipitated chalk, (b) one of the test-tubes containing powdered clay, and observe the filtrate in each case.

What effect has filtering on the turbidity of the mixture in each case?

Which filtrate appears clearer? Account for your observation.

4. Allow the remaining test-tubes to stand until the next laboratory period, and observe the appearance of each mixture again.

From your observations, do you conclude that suspensions are (a) temporary or (b) permanent? (TEXT p. 100)

EXPERIMENT 46. The formation of an emulsion.

1. Half fill a test-tube with water, and slowly pour into it approximately 2 ml. of kerosene.

Are the liquids miscible?

Account for the relative positions of the liquids.

2. Place a thumb over the mouth of the test-tube, and shake vigorously a number of times. Then place the test-tube upright in a test-tube rack, and observe any changes that take place during an interval of one minute.

Are the liquids miscible after shaking?

Account for the appearance of the liquids immediately after shaking.

Account for the appearance after the test-tube has stood for a time.

3. Using another test-tube, repeat Procedure 1. To this test-tube, add 2 ml. of soap solution, and shake vigorously a number of times.

How does the appearance of the contents of this test-tube compare with the contents of the tube used in Procedure 2, (a) immediately after shaking, (b) after standing for some time?

QUESTIONS

1. Why is the emulsion, prepared in Procedure 2, known as a *temporary emulsion*?
2. Why is the emulsion, prepared in Procedure 3, known as a *permanent emulsion*?
3. Why is soap, as used in this experiment, known as an *emulsifying agent*?
4. What is the chief distinction between a *suspension* and an *emulsion*?
5. In what respect is whole milk an emulsion?
6. In what respect is whole milk a suspension?
7. Is tomato juice a suspension or an emulsion? Explain.

UNIT 7

Two Fundamental Laws

EXPERIMENT 47. The Law of Conservation of Mass.

In this experiment a chemical reaction takes place in an enclosed space, so that the reacting substances and the products of the reaction can be weighed. The class should be divided into sections, each section being assigned the investigation of one reaction only. The substances used should be examined as *solids* before their concentrated water solutions are distributed. A class discussion should follow, involving all the reactions investigated by the various sections.

The following *pairs of solutions* are suggested: sodium chloride and silver nitrate; lead nitrate and potassium iodide; cupric sulphate and sodium hydroxide; ferric chloride and potassium hydroxide; sodium sulphate and barium chloride.

Solutions of sodium chloride and silver nitrate are used in the description that follows. The same experimental procedures apply when different pairs of substances are used.

1. The apparatus used consists of a flask and a small test-tube which will rest inside the flask, but cannot fall over (Fig. 46).

Half fill the test-tube with the silver nitrate solution. Pour about 25 ml. of sodium chloride solution into the flask, and then carefully insert the test-tube and contents in an upright position. Insert a rubber stopper firmly, and weigh the stoppered flask and its contents (Fig. 47A).

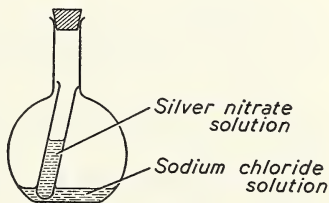


Fig. 46.

2. Leave the weights in place on the scale pan of the balance used. Remove the stoppered flask from the balance; tilt it to mix the reagents. Immediately replace the flask on the balance pan and reweigh. Remove the flask from the balance, and examine the contents.

Describe the appearance of the solids used in making the solutions and the appearance of their water solutions.

What evidence is there that a chemical change has taken place?

Name the new substances that are formed. Identify the precipitate formed.

Define a precipitate.

Has there been any change in weight as a result of the reaction?

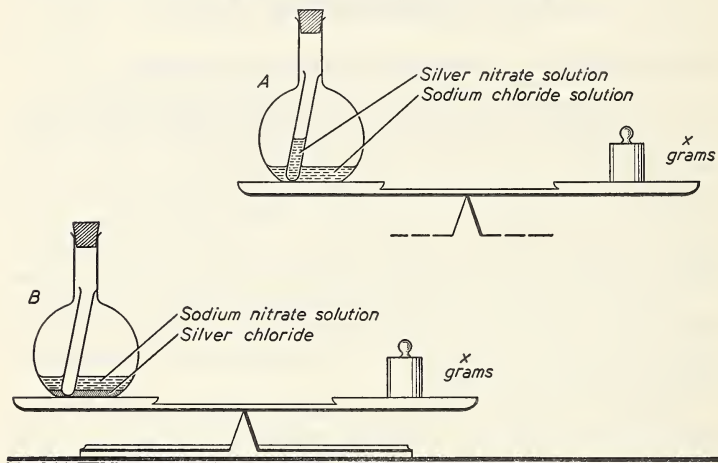


Fig. 47. The Law of Conservation of Mass.

3. Copy from the blackboard the results of the experiments carried out by other members of the class, and study these results.

What two results are common to all these experiments?

4. Assume that if many thousands of chemical reactions were studied, results similar to those found in the class experiments would apply.

State a general principle or law that would incorporate all chemical reactions. Name this law.

EXPERIMENT 48. The Law of Definite Proportions.

1. In Experiment 8 it was found that magnesium (or copper) increased in weight when heated in air. Re-examine the observations made in this experiment, and calculate the percentage composition by weight of the magnesium oxide (or copper oxide) synthesized. Compare your results with those of the class who were working on the same synthesis. (These results should be listed on the blackboard.)

Why is the reaction known as a synthesis?

What is the percentage of (a) magnesium, (b) oxygen, in magnesium oxide?

How does your result compare with those obtained by other members of the class?

Within the limits of experimental error, what may be stated about the class results?

2. Dry a clean Pyrex test-tube by heating it in a Bunsen flame. Allow it to cool, and weigh it accurately. Using a long-stemmed funnel, add about 1 gm. of mercuric oxide, making sure that none of the solid touches the sides of the tube. Remove the funnel and weigh again (Fig. 48).

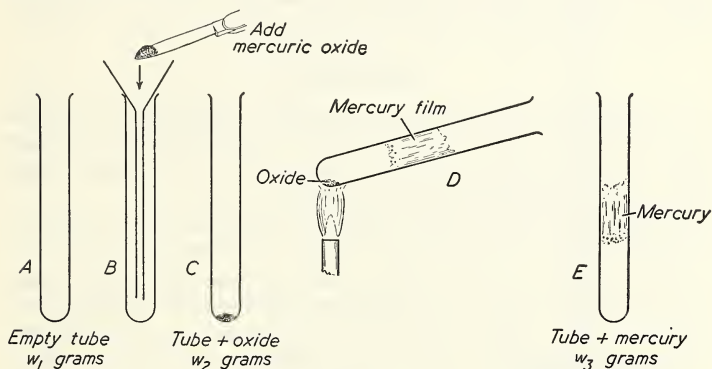


Fig. 48. The Law of Definite Proportions. The percentage composition by weight of mercuric oxide may be determined from this experiment.

Hold the tube in a test-tube holder at an angle of about 15° to the horizontal, and heat strongly at the closed end until the mercuric oxide is completely decomposed. (**Keep the flame directly beneath the mercuric oxide and avoid overheating the sides of the tube.**) Allow the tube to cool to room temperature, and weigh the tube and its contents again.

Record your observations as follows:

1. Weight of empty tube (A)..... gm.
2. Weight of tube and mercuric oxide (C)..... gm.
3. Weight of tube and mercury film (E)..... gm.

Calculate:

- (a) Weight of mercuric oxide used ($w_2 - w_1$)..... gm.
- (b) Weight of mercury obtained ($w_3 - w_1$)..... gm.
- (c) Weight of oxygen given off ($a - b$)..... gm.

Calculate the percentage composition of mercuric oxide.

What is the authority for saying that the weight of oxygen produced is equal to the weight of the mercuric oxide used minus the weight of the mercury obtained?

3. Record (from the blackboard) the results obtained by other members of the class.

How do the class results compare? Allowing for experimental error, what might be stated about these results?

Make a statement about the ratio by weight of the constituents of mercuric oxide.

List all the sources of error in this experiment.

Why is it not advisable to attempt to decompose the mercuric oxide that forms again on the sides of the tube just below the film of mercury? What caused the formation of the oxide at this place?

QUESTIONS

1. If time had permitted, you would have been asked to heat the empty test-tube used in the mercuric oxide experiment *to constant weight*, before adding the oxide. How would you do this? What effect would this have had on your result?

2. Why is the experiment involving magnesium oxide called a *synthesis*, and that with mercuric oxide an *analysis*?

3. In addition to magnesium oxide and mercuric oxide, many compounds have been both synthesized and analysed, and in every case, the proportion by weight of the constituents that make up the compound has shown no variation. State the law that enunciates this principle. What is this law called?

UNIT 8

The Gas Laws

EXPERIMENT 49. (*Demonstration*) The measurement of air pressure.

1. Select a thick-walled glass tube about three feet in length and closed at one end. Carefully fill this tube with mercury. (During the filling process, it is advisable to *tap* the tube gently at intervals to prevent the retention of trapped air bubbles.) When it is filled, hold a finger firmly over the open end of the tube, and invert it. Keep the finger in position until the end of the tube is submerged beneath the surface of mercury in a dish (cistern). Support the tube in a vertical position, and then remove the finger (Fig. 49).

State what is observed when the finger is removed.

Why does the mercury oscillate for a time? Why does it eventually come to rest?

What is present in the space above the mercury in the tube?

2. Clamp a metre stick so that the zero mark just touches the top of the mercury meniscus in the cistern. Read the length of the mercury column (in both millimetres and inches) by reading from the top of the mercury meniscus in the cistern to the top of the mercury meniscus in the tube, and record these measurements.

What is the height of the mercury column in (a) millimetres, (b) inches?

Explain how units of length may be used to measure atmospheric pressure. (TEXT p. 120)

Why is such a tube called a Torricellian tube? Why is it also called a barometer?

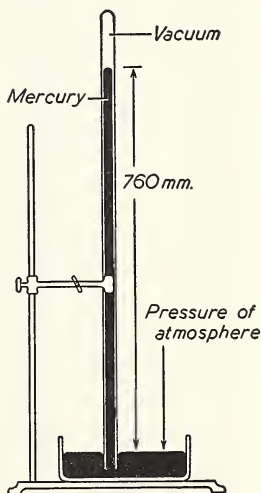


Fig. 49. A simple mercury barometer. Any change in the height of the column of mercury indicates a change in the pressure exerted by the atmosphere.

3. Examine the barometer at different times every day for a period of one week, and record the heights of the mercury column.

Explain the variations noted.

EXPERIMENT 50. (*Demonstration*) To determine the relation between the volume of a gas and the pressure to which it is subjected. (Boyle's Law)

1. Use a glass tube at least 4 mm. in diameter and 96 cm. long. The tube is bent in a U-form so that one arm is much shorter than the other. The shorter arm is sealed. Such a tube is often called a Boyle's Law tube. (If such a tube is not available, two straight tubes, one short and sealed at one end and the other longer and open at both ends, may be connected by a heavy rubber tube and substituted for the Boyle's Law tube.)

Support the tube vertically on a stand.

2. Pour mercury into the tube until it *just fills the bend* (Fig. 50A). Then manipulate the tube back and forth until the mercury reaches the same level in both arms. This is done so that the pressure of the enclosed air in the shorter arm of the tube will be the same as the pressure of the atmosphere.

Clamp a metre stick to the stand in a vertical position, so placed that the height of both arms of the tube can be measured. Record the position of the mercury level (*in centimetres*) in each arm of the tube, and the volume (*represented by the height in centimetres*) of the air enclosed in the shorter arm.

Add mercury to the tube until the level in the longer arm is about 15 cm. above the level in the shorter arm, and record the height of the mercury levels in both arms, as indicated by the readings on the metre stick (Fig. 50B). Record also the volume (*height in centimetres*) of the enclosed air. Increase the height of the mercury column in the longer arm by about 15 cm., by adding more mercury to the tube. Observe and record the positions of the mercury levels in each arm of the tube. Record the volume of the enclosed air. Continue to add more mercury at intervals until the longer arm is nearly filled (Fig. 50C). Each time mercury is added, take a reading of the mercury levels in both the longer and shorter arms of the tube, and record the volume of the enclosed air.

3. Read a mercury barometer (*in centimetres*), and record the reading. Let this be represented by A.P. (atmospheric pressure).

4. Let V , V_1 , V_2 , V_3 , etc., represent the volumes of air enclosed in the shorter arm of the tube at successive observations as indicated by the lengths of the air column.

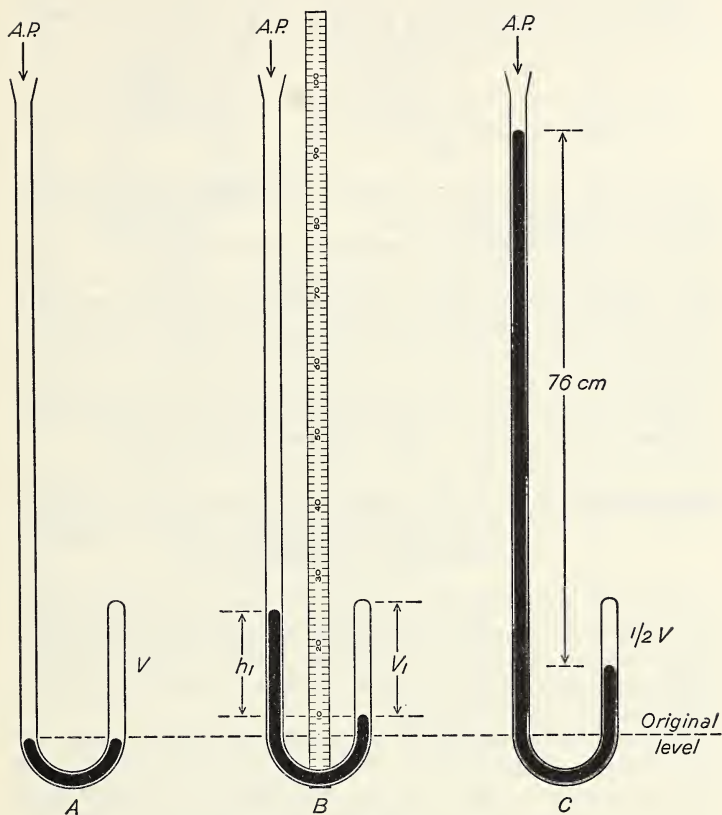


Fig. 50. A Boyle's Law Tube. In *A*, a volume of gas, V , is enclosed in the short arm of the bent tube. Because the mercury is at the same level in both arms, the pressure exerted by the enclosed gas is equal to atmospheric pressure, in this case assumed to be 76 cm. (29 in.). In *B*, mercury has been added to the open arm of the tube, showing a difference in level of h_1 cm. of mercury. This has caused a reduction in the volume of the enclosed gas, represented by V_1 . In *C*, the mercury in the long arm is 76 cm. higher than in the short arm, and the pressure exerted on the enclosed gas is 2×76 cm. The volume of the enclosed gas is half its original volume, and is represented by $\frac{1}{2}V$.

Let h_1, h_2, h_3 , etc., represent the *differences* in mercury levels in the two arms of the tube at successive observations.

Complete the following table:

VOLUME OF ENCLOSED AIR	PRESSURE EXERTED	PRODUCT OF VOLUME AND PRESSURE
V	$A.P.$	$V \times A.P.$
V_1	$A.P. + h_1$	$V_1 \times (A.P. + h_1)$
V_2	$A.P. + h_2$	$V_2 \times (A.P. + h_2)$
V_3	$A.P. + h_3$	$V_3 \times (A.P. + h_3)$
etc.	etc.	etc.

What can be said (within the limits of experimental error) about the products of the volume and pressure?

What is the percentage error (variation) in the experiment?

5. If the above experiment were carried out using *any* gas other than air, corresponding results would be obtained provided moderate pressures were used. Knowing this, and assuming that no temperature change occurs during the experiment, and that pressures do not exceed 3 atmospheres:

Make a generalization that involves volume, pressure, and temperature, for a given mass of any gas. What is this generalization called?

EXPERIMENT 51. (*Demonstration*) To determine the coefficient of expansion (volumetric) of air at a definite temperature. (Charles's Law)

1. The apparatus (Fig. 51) consists of a clean, dry, 100 millilitre flask, a 200 millilitre beaker, thermometer, retort stand, gauze, burner, one-hole stopper, short pieces of glass and rubber tubing, clamp, and a sheet-metal cover with a semi-circular hole of suitable size to accommodate the neck of the flask.

Insert the short glass tube in the stopper and attach the piece of rubber tubing, making sure that there are no leaks. Insert the rubber stopper in the flask, and place the flask in the beaker which contains water to a depth of about one inch. Place the clamp loosely over the rubber tubing. Cover the beaker with the sheet metal and, heat until the water has been boiling for about five minutes (Fig. 51A).

2. Tighten the clamp, thereby enclosing a volume of heated air. Invert the flask in cold water in the sink or a pneumatic trough. Make sure that the stopper and tube are kept covered with water. Remove the clamp, and hold the flask under water to cool the air until no more water enters it (Fig. 51B).

Equalize the levels of the water inside and outside the flask to ensure that the air pressure inside is balanced by the atmospheric pressure outside (Fig. 51C). Clamp the tube under water at the same place it was clamped before. Remove the flask from the sink and

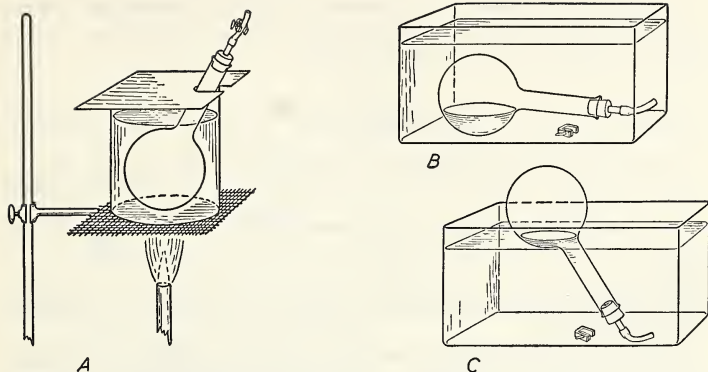


Fig. 51. Determining the coefficient of expansion (volumetric) of air at a particular temperature.

turn it right side up on the table. Remove the clamp, and loosen the stopper in order to allow any water that may be trapped in the tubing to run into the flask. Remove the stopper, and take the temperature inside the flask close to the contained water. Record this temperature in degrees centigrade.

Pour the water from the flask into a graduated cylinder, and record the volume of the water that entered the flask.

Fill the flask with water, and insert the stopper to the same depth as before. Clamp the rubber tubing at the same position as before, and flip away any water from the tubing above the clamp. Remove the clamp, and then the stopper, allowing any water contained in the glass and rubber tubings to fall back into the flask. Measure the volume of water in the flask.

Record your observations as follows:

1. Temperature of the boiling water (approximately)..... 100°C .
2. Temperature of the air after cooling (resulting temp.)..... $^{\circ}\text{C}$.
3. Volume of water that entered flask.....ml.
4. Volume of water needed to fill flask.....ml.

Why is 100°C . merely an approximation of the temperature of the boiling water?

Explain why the volume of water needed to fill the flask (Observation 4) represents the volume of air at 100°C . that was cooled during the experiment.

Explain why the volume of water that entered the flask (Observation 3) represents the shrinkage in volume of the air at 100°C . when cooled to the resulting temperature.

Calculate the coefficient of expansion of air at the temperature (Observation 2) recorded above. (For assistance in these calculations consult pages 124 and 125 of the TEXT.)

QUESTIONS

1. Given that the coefficient of expansion of air at 0°C. is $\frac{1}{273}$, explain why the coefficient of expansion of air at 10°C. is $\frac{1}{283}$. (TEXT p. 126)
2. Explain how Lord Kelvin developed the *Kelvin* or *absolute* temperature scale.
3. You are told that all gases have the same coefficient of expansion for any given temperature, and you have noted that no pressure changes occurred during the above experiment. Make a generalization that involves volume, absolute temperature, and pressure, for a given mass of any gas. What is this generalization called?

UNIT 9

Terms Involved In Chemical Combination

EXPERIMENT 52. To find the gram-equivalent weight of magnesium.

1. Cut off an 8 inch strip from a roll of magnesium ribbon. Clean the metal by pulling it through a folded piece of emery paper or fine sandpaper. Roll it into a compact coil by twisting it around a match or a piece of glass tubing. The coil must be compact enough to slide easily into a gas-measuring tube (capacity 200 ml.). Weigh the magnesium coil carefully, and record its weight. Attach a piece of thread to one end of the coil, leaving about three feet of thread free.

A note to the teacher: Much time can be saved if the teacher cleans a strip of magnesium ribbon of such length that it will provide enough 8 inch strips required for all his classes. If x strips are required, clean and weigh a strip exactly $8x$ inches in length in advance of the lesson, and record its weight. Then, using a ruler and a razor blade, cut it carefully into separate 8 inch strips. Knowing the weight of the long strip, the weight of each 8 inch strip may be *calculated* by the students.

If 200 millilitre gas-measuring tubes are not available in your laboratory, 100 millilitre graduated cylinders may be used instead.

A 4 inch strip of magnesium should be used if 100 millilitre cylinders are substituted.

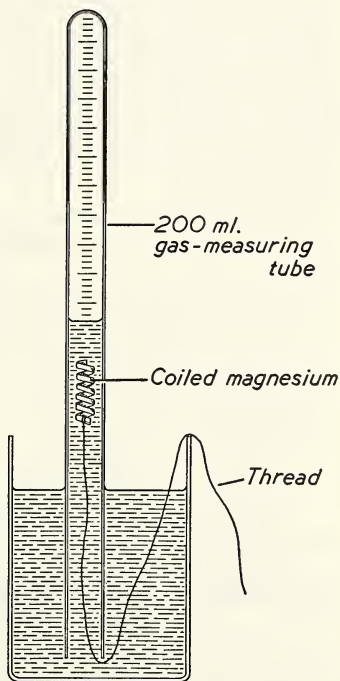


Fig. 52. Determining the gram-equivalent weight of magnesium.

2. Fill a deep battery jar or deep sink three-quarters full of water. Place about 20 ml. of concentrated hydrochloric acid in the gas-measuring tube (or 10 ml. if a 100 millilitre cylinder is used), and then slowly fill it with water by pouring along the side of the *sloping* tube to avoid, as much as possible, mixing the acid and the water. Place a piece of paper or the palm of the hand over the open end of the tube, and invert it in the water in the jar (or sink). Remove the paper or hand and *immediately* push the weighed magnesium coil up into the tube (Fig. 52). Support it there with a finger until it begins to rise in the tube, and let it continue to rise to the top of the solution. If the magnesium should stick to the side of the tube as the action proceeds, a pull on the thread will return it to the acid solution.

Why does the magnesium rise in the tube?

What gas is being collected? What is the source of this gas?

Write a word equation for the action of magnesium on hydrochloric acid.

3. When all action has ceased, cool the tube by pouring water from the jar or sink over it for five minutes. Adjust the measuring tube so that the surface of the liquid inside it is at the same level as the surface of the liquid in the jar. Carefully read and record the volume of the hydrogen collected.

Why was the water from the jar poured over the tube?

Why were the water levels adjusted before reading the final volume of the gas?

4. Read and record the temperature of the liquid in the jar (or sink).

5. Read and record the barometric pressure of the atmosphere.

6. Tabulate your data as follows:

1. Weight of magnesium used. gm.
2. Volume of hydrogen collected. ml.
3. Temperature of hydrogen. °C.
4. Barometric pressure. mm.
5. Pressure due to water vapour at the temperature recorded
in (3)—(See table p. 186, Appendix). mm.
6. Pressure exerted by dry hydrogen (4—5). mm.

NOTE: The pressure exerted by the atmosphere (as shown by the barometric reading) is equal to the pressure exerted by the hydrogen plus the pressure exerted by water vapour which is present with the hydrogen. For example, if the barometric reading is 754 mm. and the recorded temperature is 20°C., the water vapour pressure (as shown in the table p. 186) is 17.5 mm., and the corrected pressure exerted by dry hydrogen is $(754 - 17.5) = 736.5$ mm.

7. The volume of hydrogen at ____°C. and at the corrected pressure of ____ mm. equals ____ ml. *Calculate* the volume this gas would occupy at S.T.P.

8. Given that 22,400 ml. of hydrogen at S.T.P. weigh 2 gm., *calculate* the weight in grams of the hydrogen collected.

9. *Calculate* the gram-equivalent weight of magnesium by finding the weight of magnesium required to liberate 1 gm. of hydrogen from the acid.

10. List the class results and take the class average.

What are the possible sources of error in this experiment.

EXPERIMENT 53. (*Demonstration*) To find the gram-molecular weight of oxygen.

1. Dry a tablespoonful of potassium chlorate by gently warming it in an evaporating dish.

2. On a piece of glazed paper, mix about 6 gm. of this dry potassium chlorate with one-third its volume of manganese dioxide, and then transfer the mixture to a dry Pyrex test-tube. Insert a loose plug of glass wool above the mixture, and then insert a stopper containing a short piece of glass tubing. Weigh this generator and its contents.

3. Arrange the apparatus as shown in Figure 53. A deep pneu-

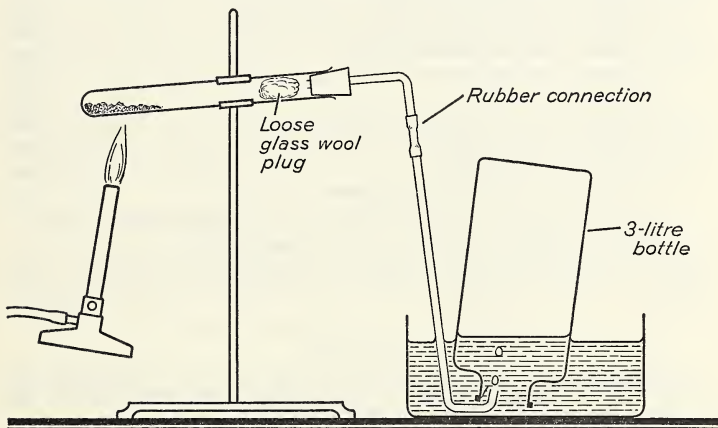


Fig. 53. Determining the gram-molecular weight of oxygen.

matic trough should be used, and the capacity of the collecting bottle should be about 2 to 3 litres. (The glass bottles in which acids are shipped to schools make good collecting bottles in this experiment.)

Fill the collecting bottle with water, and invert it in the pneumatic trough. Place the end of the delivery tube inside the bottle, and

carefully heat the oxygen generator. The heating should be controlled so that the oxygen is evolved slowly. You should be able to count the bubbles as they rise in the collecting bottle. Continue heating the mixture until the collecting bottle is about three-quarters filled with gas. Stop heating, and when the bubbles stop rising in the collecting bottle, **immediately** remove the delivery tube from the pneumatic trough.

4. Adjust the collecting bottle until the level of the water inside is the same as the level of the water in the trough. Maintaining this level, cover the mouth of the bottle with a glass plate, and then turn the bottle right side up and place it on the desk top. Remove the cover, and using a graduated cylinder, add sufficient water to fill the collecting bottle. Keep a record of the volume of water required to fill the bottle. This represents the volume of the oxygen produced.

Explain.

5. Record the temperature of the water in the pneumatic trough.

6. Record the barometric pressure of the atmosphere.

7. Reweigh the oxygen generator when it has cooled.

8. Record your data as follows:

1. Weight of generator and contents before heating.....gm.
2. Weight of generator and contents after heating.....gm.
3. Weight of oxygen liberated (1—2).....gm.
4. Temperature of oxygen (temp. of water in trough).....°C.
5. Barometer reading.....mm.
6. Pressure due to water vapour (See table p. 186).....mm.
7. Pressure of dry oxygen (5—6).....mm.
8. Volume of gas collected.....ml.

9. Make the following calculations:

(a) The volume that the collected oxygen would occupy at S.T.P.

(b) The weight of 22.4 litres of oxygen at S.T.P.

10. The gram-molecular weight of oxygen as determined by this experiment is ____ gm.

QUESTIONS

1. What was the purpose of using a loose glass wool plug in this experiment?

2. Why was the generator heated carefully?

3. Why were the first bubbles, which obviously were air, collected in this experiment, rather than being discarded as they were in the laboratory preparation and collection of oxygen? Is this a serious source of error? Explain.

4. Why was it necessary to measure the temperature and pressure of the gas collected,

5. List all sources of error in the above experiment.
6. What is the accepted value for the gram-molecular weight of oxygen? What is its molecular weight? (TEXT pp. 145 and 150)
7. What is the percentage error in this experiment.

EXPERIMENT 54. To find the gram-molecular weight of carbon dioxide.

1. Record the room temperature and the barometer reading.
2. Place a rubber stopper in a clean, dry, 2 litre flask, and weigh the stoppered flask. Record the weight.

Given that 1 litre of air at S.T.P. weighs 1.29 grams, calculate the weight of 2 litres of air at room temperature and at a pressure indicated by the barometer reading.

What would the stoppered flask weigh if all the air were removed?

3. Connect a piece of rubber tubing to a cylinder of carbon dioxide, and insert a piece of glass tubing about a foot long into the rubber tubing. Remove the rubber stopper from the 2 litre flask, and insert the glass tubing into the flask so that it almost touches the bottom. Open the valve on the cylinder of carbon dioxide, and allow carbon dioxide to escape, under pressure, into the flask for about 10 seconds. By this time, all the air within the flask will be replaced with carbon dioxide. Insert the stopper, and again weigh the flask. Record this weight.

What is the weight of the stoppered flask filled with carbon dioxide?

Calculate the weight of two litres of carbon dioxide at room temperature and pressure. What would this volume be at S.T.P.?

Calculate the weight of 22.4 litres of carbon dioxide at S.T.P.

What is the gram-molecular weight of carbon dioxide as determined by this experiment? What is the correct gram-molecular weight of carbon dioxide? (TEXT p. 145) What is the percentage error?

What is the molecular weight of carbon dioxide?

EXPERIMENT 55. To find the valence of a metal by the displacement of hydrogen from an acid.

1. The experimental method is the same as that used in Experiment 52. The average of the class results in Experiment 52 may be taken as a starting point for determining the valence of magnesium, or the experiment may be repeated. The gram-atomic weight of magnesium is taken as 24.32 grams.

What is meant by saying that the gram-atomic weight of magnesium is 24.32 grams?

What weight of magnesium was used to displace the hydrogen from the dilute hydrochloric acid?

What was the calculated volume of hydrogen (collected at S.T.P.) displaced by this weight of magnesium?

2. Calculate (to the nearest whole number) the weight in grams of hydrogen that would have been liberated from hydrochloric acid if 24.32 grams of magnesium had been used. This number (i.e., the number of grams of hydrogen displaced) is the valence of magnesium.

What is the valence of magnesium? Define valence.

Write the word equation for the chemical reaction that takes place when magnesium is added to hydrochloric acid. How many atoms of hydrogen are liberated from hydrochloric acid by one atom of magnesium?

3. Using a hydrogen generator (with delivery tube, etc.), a weighed amount of zinc, a large gas bottle, and a pneumatic trough, tell how you would find the valence of zinc. You are told that the gram-atomic weight of zinc is 65.38 grams.

If the teacher approves your method, proceed to find the valence of zinc.

UNIT 10

Ionization, Acids, Bases, and Salts

EXPERIMENT 56. (*Demonstration*) The conductivity of various solutions.

1. Fit two copper electrodes to a piece of Bakelite or other insulating material in such a way that connecting wires may be attached which will lead to a switch, a lamp, (a non-frosted glass bulb showing the filament if possible), and a source of electricity, as shown in Figure 54.

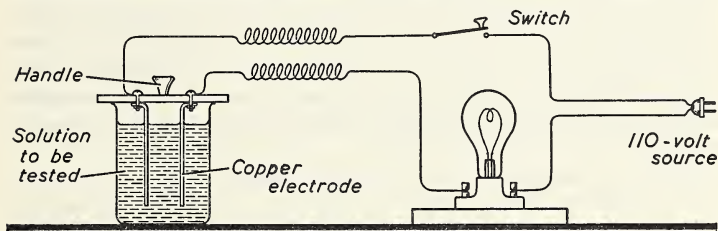


Fig. 54. Testing the conductivity of a solution.

When the apparatus has been properly assembled, plug it into a 110-volt source of electricity and close the switch. Using an insulating material such as a glass rod, complete the circuit by bending the two copper electrodes until they touch, and then separate them to their original positions.

Describe and account for the appearance of the light bulb.

Distinguish between the terms "open circuit" and "closed circuit".

2. Nearly fill a clean, 250 millilitre beaker with distilled water. Surround the copper electrodes with the distilled water, close the switch, and observe the filament of the lamp and also the electrodes in the distilled water. **Open the switch.**

Does the filament glow?

Is there any indication that electricity is flowing through the circuit?

Does distilled water conduct an electric current?

3. Keeping the electrodes surrounded with distilled water as in Procedure 2, close the switch, and add concentrated sulphuric acid,

one drop at a time with constant stirring, until approximately 2 ml. of acid have been added. Observe the filament of the lamp closely after each addition of acid, and also observe the electrodes in the solution. **Open the switch.**

Does the filament glow?

Is there any indication that electricity is flowing through the circuit?

Does a water solution of sulphuric acid conduct an electric current?

Is any noticeable change taking place within the beaker?

4. In separate 250 millilitre beakers, prepare solutions containing approximately 2 gm. of each of the following substances in distilled water: sodium hydroxide, ammonium hydroxide, sodium chloride, concentrated acetic acid, potassium nitrate, cane sugar, ethyl alcohol, and glycerine. In each case the total volume of solution should equal that used in Procedure 2. **NOTE: To determine relative conductivities accurately, solutions of equivalent concentrations must be used.**

Using a beaker containing distilled water, rinse the copper electrodes thoroughly to remove any traces of sulphuric acid remaining on them from Procedure 3. Close the switch, and surround the copper electrodes with the contents of each separate beaker in turn, **rinsing the electrodes thoroughly in distilled water before each new solution is tested.**

5. Observe the brightness of the lamp filament for each solution, and compare the effect with that produced in Procedures 1 and 2.

List the solutions that apparently do not conduct an electric current.

Arrange the substances whose solutions conduct an electric current in the order of their conductivity, as indicated by the brightness of the lamp during the experiments.

Does the fact that the filament does not glow indicate that no electricity is passing through the circuit? Explain.

QUESTIONS

1. What three main classes of substances make good conductors of electricity when dissolved in water?

2. Use Arrhenius' theory of ionization to explain why electrolytes conduct an electric current. (TEXT p. 248)

3. By reference to the theory of ionization (TEXT p. 249), explain the electrolysis of water (Experiment 35).

EXPERIMENT 57. The properties of acids.

1. Into separate clean test-tubes half-filled with water, put 5 drops of each of the following concentrated acids: hydrochloric, sulphuric, acetic, nitric, and phosphoric.

2. Put a drop of each *dilute* solution on the end of a clean glass stirring rod, and taste each drop in turn.

3. Rub a drop of each dilute solution between the thumb and the forefinger. Compare the "feel" to that of water.

4. Dip a clean glass rod into each solution, and touch it to both red and blue litmus paper (or use a single piece of neutral litmus paper).

5. Divide each of the acid solutions into four equal parts by pouring each solution into three additional test-tubes, and complete the following tests for each acid. (If only three additional test-tubes are available, rinse these three thoroughly before proceeding to the next acid.)

6. Add a few drops of pink phenolphthalein to one portion, and note any colour change.

7. Add a few drops of neutral (green) bromthymol blue solution to another portion, and note any colour change.

8. Add a pinch of baking soda to another portion, and note any reaction. Dip a piece of glass tubing in limewater and hold it vertically, so that by placing a finger over its upper end, a few drops of limewater are retained in the tube. Lower the tubing into the test-tube, and observe any change in the limewater.

9. Add a pinch of magnesium powder to the final portion, and note any reaction. Lower a blazing splint into the test-tube, and describe the result.

Complete the following table in your notebook, and include *all* the observations noted in the seven tests made with each acid. Identify the gas produced in all cases where a gaseous product is noted.

	HYDRO- CHLORIC	SULPHURIC	ACETIC	NITRIC	PHOSPHORIC
Formula					
Taste					
Feel					
Action on litmus					
Action on phenolphthalein					
Action on neutral bromthymol blue					
Action on baking soda					
Action on magnesium					

List all of the properties that are common to each of the acids tested.

EXPERIMENT 58. The properties of bases.

1. Pour into separate clean test-tubes, to the depth of one-half inch, concentrated solutions of each of the following bases: sodium hydroxide, potassium hydroxide, ammonium hydroxide, and calcium hydroxide.

2. Add water to each tube until three-quarters full, and then mix the contents of each tube thoroughly by stirring with a clean glass rod.

3. Carry out the tests suggested for *acids* (Experiment 57.) Test each solution for (a) taste, (b) "feel" (rinse the fingers after each test), (c) effect on litmus, (d) effect on *colourless* phenolphthalein, (e) effect on neutral (green) bromthymol blue, (f) effect on baking soda, (g) effect on magnesium.

Construct a table for bases, similar to that used for acids, and tabulate all the observations made.

List all the properties that are common to each of the bases tested.

EXPERIMENT 59. Neutralization.

1. Pour 5 ml. of a dilute sodium hydroxide solution into a clean evaporating dish, and add 2 drops of colourless phenolphthalein solution. Note the colour produced.

2. Add dilute hydrochloric acid to the sodium hydroxide solution with a medicine-dropper, drop by drop, stirring after each addition, until the colour just fades. Be careful not to add an excess of the acid.

Is the resulting solution basic? Is it acidic? Explain.

3. Place the evaporating dish on a piece of wire gauze resting on a ring clamped to a retort stand. Evaporate the solution to dryness. If the flame is removed just before the solution reaches dryness, the heat retained by the evaporating dish will dry the residue. Allow the dish to cool, and examine and taste the residue.

Describe the appearance of the residue.

What is the taste of the residue? Name the residue.

Write an equation to represent the reaction which occurs when hydrochloric acid is added to sodium hydroxide solution. Interpret this equation in terms of relative weights.

What is the name used to indicate this type of reaction? What is the significance of this term?

EXPERIMENT 60. Titration.

When a neutralization is carried out in a quantitative way, the procedure is called titration.

The instructor has prepared two solutions. One solution is sodium hydroxide and contains 40 gm. of sodium hydroxide per litre of solution. The other solution contains hydrochloric acid of a concentration somewhere between 40 gm. and 50 gm. of hydrogen chloride per litre of solution. The exact concentration is known to your

instructor. Your problem is to determine the concentration of the hydrochloric acid in grams of hydrogen chloride per litre of solution.

1. Consult *Technique 1*, and study the structure, the calibration, and the use, of a burette. Rinse two clean burettes with distilled water. (A 25 millilitre pipette may be substituted for one burette.)

2. Rinse one burette with a few millilitres of the sodium hydroxide solution. Clamp it to a retort stand, fill it slightly above the zero mark with the sodium hydroxide solution, and adjust the level of the liquid to the zero mark (or any other recorded position). Discard the surplus base. Rinse the second burette with the hydrochloric acid that is to be tested, clamp it to a retort stand, and fill it slightly above the zero mark with the acid. Adjust the level of the liquid to the zero mark (or any other recorded position). Discard the surplus acid.

3. Place a clean beaker below the burette containing the sodium hydroxide solution, and run in exactly 25 ml. of solution (or add the solution to the beaker from a 25 ml. pipette). Add 2 drops of colourless phenolphthalein solution, and then place the beaker on a piece of white paper below the burette containing the hydrochloric acid.

Allow 2 ml. of the hydrochloric acid to run into the beaker, and then stir with a clean glass rod. Continue to add the acid in 2 millilitre volumes, stirring after each addition, until the solution becomes colourless. This is a *test run* to get the *approximate volume* of acid solution necessary for the neutralization. Note the volume of acid required and discard the liquid in the beaker.

4. Refill both burettes, and adjust the level of the liquid in each one to the zero mark (or any other recorded position), as in Procedure 2. Wash and dry the beaker used earlier, and repeat the titration. This time, run the acid into the sodium hydroxide solution rapidly, until a volume of acid approximately 2 ml. less than that used in the test run has been added. Then continue to add the acid drop by drop, stirring after each drop, until the colour just changes.

Read the burette, and determine the volume of hydrochloric acid required to neutralize the base.

5. Repeat the titration and find the *average* volume of hydrochloric acid used.

What was the average volume of the hydrochloric acid solution used in the titration?

Write the equation for the reaction.

What weight of hydrogen chloride is necessary to neutralize 40 gm. of sodium hydroxide? (This information is indicated by the equation for this reaction.)

If a litre of the sodium hydroxide solution contains 40 gm. of sodium hydroxide, what weight of sodium hydroxide is present in the 25 ml. of solution neutralized in the titration?

What weight of hydrogen chloride (HCl) would be needed to neutralize this weight of sodium hydroxide (i.e., the weight of the sodium hydroxide in the 25 ml. of the solution used)?

What volume of acid solution (measured in millilitres) contains this weight of hydrogen chloride?

Calculate the weight of hydrogen chloride that would be present in a litre of the hydrochloric acid solution used in the titration.

UNIT 11

Carbon and its Compounds

EXPERIMENT 61. Common forms of carbon.

1. Examine small samples of each of the following common forms of carbon: graphite, wood charcoal, soft coal, hard coal, coke, boneblack, lampblack, and carbon black. Attempt to mark a piece of paper with each form, and note the "feel" of each.

Describe the colour, hardness, feel, and general appearance of each form of carbon.

Name the forms which leave a mark on paper?

Which form is most commonly used as a writing material? Why?

Does graphite give the impression that it is crystalline in structure? Is it crystalline? (TEXT p. 260)

2. Heat a small sample of graphite as intensely as possible in a Bunsen flame.

What is the effect of intense heat on graphite?

What use of graphite does this suggest?

EXPERIMENT 62. Common sources of carbon.

1. In this experiment, samples of each of the following substances are separately tested: wood, paper, cotton, wool, bread, potato, carrot, corn starch, sugar, meat, and paraffin wax.

Place a thin layer of clean quartz sand in a small crucible, and on it, put a small sample of the substance to be tested. With the crucible cover on, heat strongly for five minutes. Cool the crucible, and examine the contents. For a class experiment, it is suggested that each group work with a different substance.

Describe the change in appearance of the original substance in each case. What constituent is present in all the samples tested?

2. Place a small piece of each residue from Procedure 1 on a piece of mica, and direct a hot Bunsen flame on each in turn.

Does the residue burn? Explain why burning did not occur in Procedure 1.

3. (a) Completely close the air inlet on a lighted Bunsen burner, and reduce the gas pressure until the flame is very small. Hold a

glass plate in the tip of the flame for a short time, and observe the surface of the glass.

(b) Hold a glass plate in the flame of a paraffin candle, and after a short time, observe the surface of the glass.

(c) Place 2 ml. of kerosene in an evaporating dish, and ignite it. Hold a glass plate in the flame, and after a short time, observe the surface of the glass.

What is the colour of the flame in each case?

Describe the deposit formed in all three experiments? What is this deposit? What is the explanation for its formation?

QUESTIONS

1. Account for the presence of carbon in each of the substances used in the above experiment. (TEXT p. 258)

2. Explain the importance of carbon as a constituent in each of the following: wood, bread, sugar, and paraffin wax.

EXPERIMENT 63. The destructive distillation of wood.

1. Arrange the apparatus as in Figure 55, and insert either wood shavings or splints in the horizontal test-tube. Heat the test-tube with a moderate flame, and observe all the changes that take place.

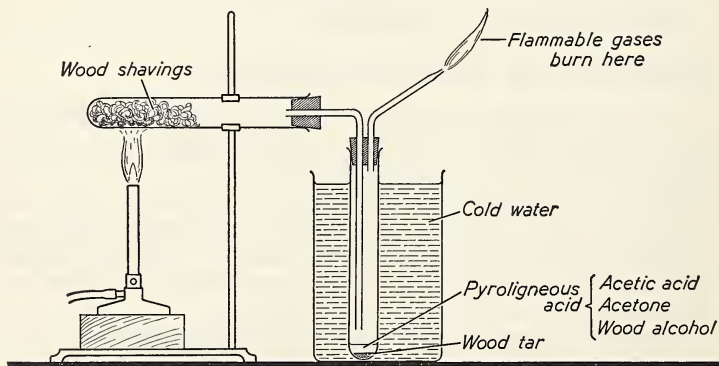


Fig. 55. The destructive distillation of wood.

When there is evidence of a substance collecting in the vertical test-tube, bring a lighted splint to the jet, and observe the colour, size, and duration of the flame.

Describe the changes that take place in the wood.

What is the nature of the deposit which forms near the mouth of the upper test-tube?

Describe the distillate which collects in the vertical test-tube.

What is the colour, size, and duration of the flame at the jet?

When does the flame go out? What is the reason for this behaviour?

What are the three distinct types of products resulting from the destructive distillation of wood?

Why did the wood not burn when heated in the test-tube?

Why is this type of distillation known as "destructive"?

2. Allow the apparatus to cool. Then remove and examine the contents of the horizontal test-tube.

Describe and name the contents. Has the wood changed throughout?

3. Note the colour, feel, and odour of the distillate. Drop small pieces of blue and red litmus paper into the liquid.

Describe the colour, feel, and odour of the distillate.

Name four substances contained in the distillate. (TEXT p. 262)

What is the effect on litmus? Which of the components of the distillate affects the litmus?

EXPERIMENT 64. The destructive distillation of coal.

1. Use the same apparatus as that used in the destructive distillation of wood (Experiment 63).

Half fill the test-tube with small pieces of soft coal. Heat the tube gently at first, and then more intensely. Hold a piece of moist neutral litmus paper at the jet and observe any change in colour. When there is evidence of a substance forming in the vertical test-tube, bring a lighted splint to the jet, and observe the colour, size, and duration of the flame.

Describe and account for the change in colour of the litmus paper.

Name and give the formula of the gas responsible for this change.

Describe the changes that take place in the coal.

Describe the distillate which collects in the vertical test-tube.

What are the three distinct types of products resulting from the destructive distillation of coal?

2. Allow the apparatus to cool, and remove and examine the contents of the horizontal test-tube.

Describe and name the residue. How does it compare in structure, appearance, and colour with the original soft coal?

3. Examine the distillate and note its colour, feel, and odour.

Describe the distillate as to colour, feel, and odour.

What visible proof do you have that there must be several different liquids present in the distillate?

EXPERIMENT 65. The properties of amorphous carbon.

1. Fill a test-tube one-quarter full of activated charcoal, and pour in a dilute solution of potassium permanganate until the tube is three-quarters full. Shake the tube for a short time, and filter the contents through a moistened filter paper placed in a funnel. Collect and examine the filtrate.

What is the colour of the original potassium permanganate solution?

What is the colour of the filtrate?

What important property of activated charcoal is illustrated by this experiment?

2. Dissolve a teaspoonful of brown sugar in 100 ml. of water in a beaker, and note the colour of the solution. Add two teaspoonfuls of boneblack to the solution, and after heating for ten minutes, filter the mixture through a moistened filter paper placed in a funnel. Collect 2 ml. of the filtrate on a watch-glass. Note the colour of this filtrate, and allow it to stand until the next laboratory lesson.

What is the colour of the original sugar solution?

What is the colour of the filtrate?

What is the nature of the residue on the watch-glass? Account for its appearance.

What important property of amorphous carbon is illustrated by this procedure?

Why is boneblack used in the refining of sugar?

3. (**Demonstration**) Pour mercury to a depth of approximately one inch into a 250 millilitre beaker. Fill a test-tube with ammonia gas, insert a piece of wood charcoal into the test-tube, and invert it in the mercury. Clamp the test-tube in a vertical position, and observe any change in the mercury level.

Describe and account for the change in the mercury level.

What important property of amorphous carbon is illustrated by this experiment?

QUESTIONS

1. Explain the action of *activated charcoal* in the canister of a gas mask. Why is activated charcoal more efficient than any other form of carbon?

2. Explain the difference between the terms *adsorption* and *absorption*.

3. Under what circumstances might a medical doctor prescribe charcoal pills?

4. What are the advantages of a charcoal dentifrice?

EXPERIMENT 66. Carbon as a reducing agent.

1. Mix intimately on a piece of paper 5 gm. of cupric oxide with an equal volume of powdered charcoal, and place the mixture in a Pyrex test-tube. Fit the test-tube with a one-hole stopper and a delivery tube which dips into 5 ml. of limewater in another test-tube. Heat the mixture strongly for about ten minutes and observe the reaction in each test-tube.

What evidence is there of a chemical reaction in the tube containing the cupric oxide?

Describe and account for the reaction that takes place in the limewater.

2. Allow the test-tube to cool. Empty the contents into an evaporating dish, and after examining carefully, add a few drops of concentrated nitric acid.

Describe and account for the colour of the material in the evaporating dish.

Describe and account for the action that takes place with nitric acid.
(TEXT p. 367)

Write an equation expressing the reaction between cupric oxide and carbon.

Explain the action of carbon as a reducing agent.

EXPERIMENT 67. Common methods of producing carbon dioxide.

1. Fill a gas bottle with oxygen (Experiment 13), and pour limewater to a depth of one-quarter inch into the bottle. Place a piece of charcoal in a deflagrating spoon, ignite the charcoal, and lower it into the bottle of oxygen. After all action has ceased, remove the spoon and shake the bottle.

Describe the behaviour of the charcoal and account for the change in the limewater.

Why is this reaction used as a test for carbon dioxide?

Write chemical equations to represent all the reactions that take place in this procedure.

2. Pour limewater into a large gas bottle to a depth of one-quarter inch and lower a burning candle into the bottle. The candle may be held with a wire or attached to the bowl of a deflagrating spoon. Remove the candle and shake the bottle.

Describe the behaviour of the candle, and account for the change in the limewater.

3. Fill a Pyrex test-tube one-third full of magnesium carbonate, and fit it with a one-hole stopper and delivery tube. Pass the delivery

tube into 5 ml. of limewater in another test-tube. Heat the test-tube containing the magnesium carbonate, and observe the limewater contained in the other tube.

Describe and account for the changes in the limewater.

Write a chemical equation for the heating of magnesium carbonate.

Repeat Procedure 3 using (a) calcium carbonate, and (b) sodium carbonate, in place of magnesium carbonate.

Does each carbonate liberate carbon dioxide when heated?

4. Into separate test-tubes, place approximately 2 gm. of each of the following substances: magnesium carbonate, calcium carbonate, sodium carbonate, and sodium bicarbonate. Into the first test-tube, pour 5 ml. of dilute hydrochloric acid, and observe the reaction. Dip a clean piece of glass tubing in limewater, hold it vertically so that a drop of lime-water collects at the tip, and bring the drop to the mouth of the test-tube. Repeat with each of the remaining three test-tubes.

What gas is given off in each case?

Write a chemical equation to represent each reaction.

5. Repeat Procedure 4 using (a) dilute sulphuric acid, (b) dilute nitric acid, and (c) dilute acetic acid.

Is carbon dioxide given off in each case?

Write a chemical equation to represent each reaction that takes place.

6. Place one teaspoonful of baking powder in a 250 millilitre beaker, and add 20 ml. of water. After the reaction has subsided, lower a blazing splint into the beaker.

Describe and account for the reaction that takes place when water is added to baking powder.

Describe and account for the behaviour of the burning splint.

7. Place 10 ml. of limewater in a large test-tube. Place a clean piece of glass tubing in the limewater, and blow your breath through it until a change takes place in the limewater.

Describe and account for the change that takes place in the limewater.

8. (**Demonstration**) Place a teaspoonful of corn syrup or molasses, 50 ml. of water, and half a cake of yeast, into a 250 millilitre flask. Fit the flask with a one-hole stopper and a delivery tube which dips into 10 ml. of limewater in a test-tube. Shake the contents of the flask so as to mix them thoroughly and set the apparatus in a warm place. Observe the mixture in the flask and the limewater in the test-tube at the end of the class period, and again the following day.

What change takes place within the flask during the class period?

Is there any change in the limewater after thirty minutes? Is there a change by the second day? Account for any change observed.

QUESTIONS

1. Name six different methods of producing carbon dioxide. Which of these methods is most suitable for producing large quantities in the laboratory?

2. Account for the presence of carbon dioxide in the atmosphere. Suggest an experiment that would prove conclusively that carbon dioxide is present in very limited amounts in the atmosphere. If the teacher approves your method, try the experiment.

EXPERIMENT 68. To prepare and collect carbon dioxide and study its properties.

PREPARATION AND COLLECTION

1. Place two teaspoonfuls of marble chips in a 250 millilitre flask, and fit the flask with a thistle-tube and a delivery tube, as illustrated in Figure 56. Add dilute hydrochloric acid through the thistle-tube until the marble chips are covered and the lower end of the thistle-tube is below the level of the acid.

2. Collect one bottle of the gas by the upward displacement of air, and three bottles by the downward displacement of water.

3. Attach a funnel to the delivery tube, and allow the gas to come into contact with water in a beaker, as illustrated in Figure 56.

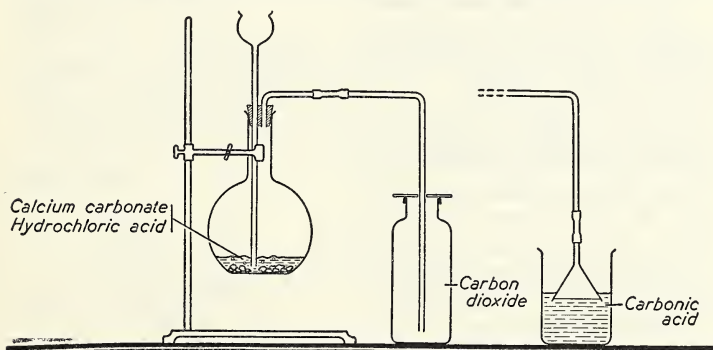


Fig. 56. The laboratory preparation and collection of carbon dioxide.

Why should the lower end of the thistle-tube be below the surface of the acid?

Which method of collection is the more satisfactory? Why?

PROPERTIES OF CARBON DIOXIDE

1. Note the colour and the odour of the gas. Lower a blazing splint slowly into a bottle of the gas.

Describe the colour and odour of carbon dioxide.

Does the gas burn? Does it support the combustion of wood?

2. Wind a 4 inch strip of magnesium ribbon into a tight spiral. Hold the metal in a pair of crucible tongs, and ignite the lower end of the coil. Slowly lower the burning magnesium into a bottle of carbon dioxide. Examine the deposit in the bottle.

Why does the magnesium continue to burn in the carbon dioxide?

Compare the flame with that of burning magnesium in air.

Describe the deposit formed, and account for the formation of any new substances.

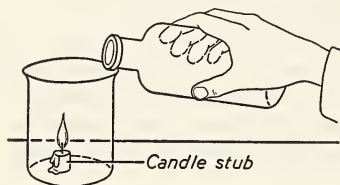


Fig. 57. Pouring carbon dioxide on a candle flame.

Write the equation to represent the reaction.

From the evidence of this experiment, is it correct to say that carbon dioxide supports combustion? Explain your answer.

3. Light a small candle and stand it upright in a beaker placed on the table. Tip a

bottle of carbon dioxide over the candle flame as if you were pouring a liquid (Fig. 57), and observe the action that takes place.

What happens to the candle flame?

What two important properties of carbon dioxide are illustrated by this experiment? Explain.

4. Add 10 ml. of water to a bottle of carbon dioxide. Cover the mouth of the bottle with the palm of your hand, and shake vigorously (Fig. 58). Add another 10 ml. of water, and repeat the procedure.

Add a few drops of neutral (green) bromthymol blue to the solution.

Account for the suction on your hand.

What is the effect of the solution on bromthymol blue?

What compound is present in solution?

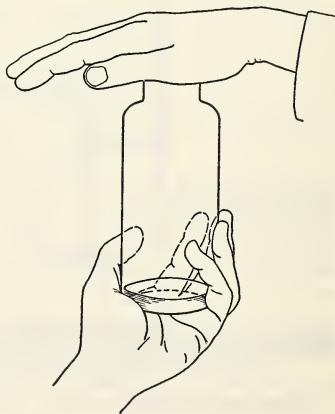


Fig. 58. Testing the solubility of a gas.

Why was the suction on the hand less noticeable when the additional 10 ml. of water was shaken with the gas?

5. If the action in the carbon dioxide generator has ceased, add either more acid or more marble chips. Fill a test-tube one-quarter full of limewater, and pass carbon dioxide from the generator through it until it becomes milky. Continue to pass carbon dioxide through the milky mixture until it clears.

Write the chemical equation that represents the reaction that produces the milky precipitate.

Explain why the precipitate disappears, and write the chemical equation that represents this reaction.

6. Drop pieces of red and blue litmus paper into the solution of carbon dioxide contained in the beaker (Fig. 56).

Describe and explain the action that takes place. Why is carbon dioxide called an acid anhydride?

7. Add 15 ml. of the carbon dioxide solution formed in the beaker to 5 ml. of limewater in a test-tube, and observe the speed with which a precipitate is formed. Note also the intensity of the precipitate.

Pour 15 ml. of solution from a bottle of *soda water* into 5 ml. of limewater in a test-tube, and observe the speed with which a precipitate is formed. Note also the intensity of the precipitate.

Which has the higher concentration of carbon dioxide, the prepared solution or the soda water? Give reasons for your conclusion.

8. Drop a small piece of carbon dioxide "snow" into a beaker containing water, and another small piece into a beaker containing limewater. Observe each action closely.

Describe carbon dioxide "snow" or "dry ice". Why is it so called?

Describe and account for the changes observed in each of the above experiments.

QUESTIONS

1. Write the equation for the laboratory preparation of carbon dioxide.

2. List all of the properties of carbon dioxide as determined in Experiment 68.

3. What three properties of carbon dioxide make it useful in extinguishing fires?

EXPERIMENT 69. (Demonstration) Common fire extinguishers.

1. **CAUTION:** Glass models of fire extinguishers should be tested by the teacher before the demonstration, in order to check on the proper concentrations and amounts of the reactants used. Stoppered bottles are extremely dangerous since the stopper may be blown out if the reaction is too violent, causing the teacher and the students to be spattered by the contents. Metal models are available from several supply houses, and are much safer.

Use a 6 oz. large-mouth bottle with vertical sides and a tight-fitting screw cap. Punch a hole one-quarter inch in diameter in the centre of the screw cap, and make sure that the paper gasket on the inside of the cap is cut well away from this quarter inch aperture. Fill the bottle three-quarters full of a solution of sodium bicarbonate. Place 5 ml. of concentrated sulphuric acid in a small test-tube, and wedge it into the bottle with a rubber stopper, so that the open end of the test-tube is well above the level of the baking soda solution (Fig. 59). Screw the cap on the bottle tightly, and invert the bottle over the sink.

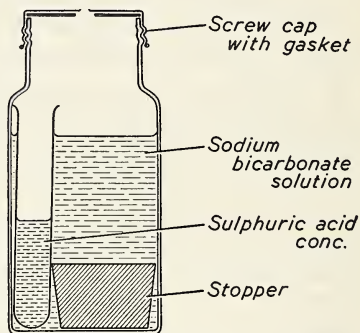


Fig. 59. A demonstration model of a soda-acid fire extinguisher.

Describe and account for the action that takes place.

Explain the action of a soda-acid fire extinguisher. (TEXT p. 270)

What class of fires may be controlled by the soda-acid fire extinguisher?

2. **(Demonstration)** Use four hydrometer jars approximately two inches in diameter and twelve inches in height. Label the jars *A*, *B*, *C*, and *D*. Pour 50 ml. of a concentrated solution of sodium bicarbonate into each jar. To jars *B* and *D*, add two teaspoonfuls of licorice extract.

To *A* and *B* in turn, add 50 ml. of dilute sulphuric acid, and observe the action that takes place.

To *C* and *D* in turn, add 50 ml. of a concentrated solution of aluminum sulphate, and observe the action that takes place.

Describe the action that takes place in jar A, and write the equation for the reaction.

Describe the action that takes place in jar B. What is the effect produced by the addition of licorice extract?

CHARACTERISTICS OF COMMON APPROVED HAND FIRE EXTINGUISHERS

Type	Plain Water*	Soda and Acid	Chemical Foam	Vaporizing Liquid		Carbon Dioxide	Dry Chemical
				Pump type	Pressure type		
Main chemicals used	Water	Sodium bicarbonate solution and sulphuric acid	Sodium bicarbonate solution and aluminum sulphate solution	Carbon tetrachloride base	Carbon tetrachloride or chlorobromomethane base	Liquefied carbon dioxide gas	Treated sodium bicarbonate powder
Expellant	Hand pump	Internal pressure from chemical reaction	Internal pressure from chemical reaction	Hand pump	Stored carbon dioxide or air pressure	Inherent pressure of liquefied gas	Carbon dioxide cartridge
Principal extinguishing agent	Water	Water	Foam	Chemical vapours	Chemical vapours	Carbon dioxide gas	Carbon dioxide gas liberated from the powder
Principal extinguishing effect	Cooling Quenching	Cooling Quenching	Blanketing Cooling	Smothering	Smothering	Smothering	Smothering
Nominal capacity	2½-5 gallons	2½ gallons	2½ gallons	1 pint 1 quart	1 quart to 2 gallons	2-20 pounds	4-30 pounds
Freezing point	32 degrees F.	32 degrees F.	32 degrees F.	50 degrees below zero F.	50 degrees below zero F.	Non-freeze	Non-freeze
Effective range	25 feet	25 feet	20-25 feet	15 feet	10-15 feet	3-8 feet	6-12 feet
Recharging material	Water	Manufacturer's charge and water	Manufacturer's charge and water	Manufacturer's extinguishing fluid only	Manufacturer's extinguishing fluid only	Manufacturer's recharge service	Manufacturer's charge and cartridge
Effective on fires (footnote)	Class A Yes	Yes	Yes	Surface only	Surface only	Surface only	Surface only
	Class B No	No	Yes**	Small fires	Yes	Yes	Yes
	Class C No	No	No	Yes	Yes	Yes	Yes

NOTE—Class A Fires—Fires in ordinary combustible materials, such as wood, paper, textile fabrics, rubbish, etc.
 Class B Fires—Fires in flammable liquids, such as gasoline, kerosene, oils and greases.
 Class C Fires—Fires in electrical equipment, such as switch boards, motors and wiring.

* Calcium chloride or other anti-freeze chemicals may be added to plain water to depress the freezing point as required. Some are corrosive.

** Not effective on fires in alcohol, acetone, carbon disulphide, lacquer thinners, ethers or esters.
 Ontario Fire Marshal's Office, Toronto, Ont.

Describe the action that takes place in jar C, and account for the colourless bubbles and the white gelatinous precipitate. Write the equation to represent the reaction.

Compare the action in jar C with that in jar A.

Describe the action that takes place in jar D. Account for the formation of the foam.

What class of fires may be controlled by a foam-type fire extinguisher?

3. Using the chart on page 119 for reference, discuss the construction, operation, and relative advantages of all types of fire extinguishers found in your school.

EXPERIMENT 70. (*Demonstration*) To prepare and collect carbon monoxide and study its properties.

CAUTION: Because of the poisonous nature of carbon monoxide, care must be taken to prevent the escape of more than small amounts of this gas into the room.

PREPARATION AND COLLECTION

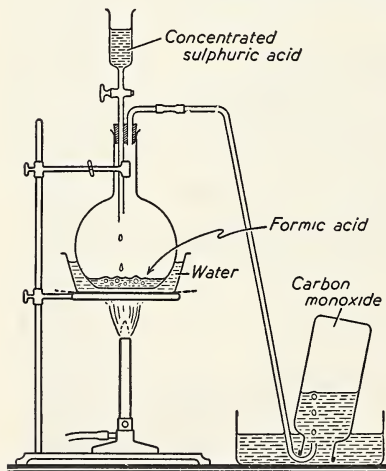


Fig. 60. The laboratory preparation and collection of carbon monoxide.

1. Arrange the apparatus as in Figure 60, and after the formic acid has been warmed slightly, let the concentrated sulphuric acid fall, one drop at a time, into the flask.

2. Collect one full bottle of the gas and one bottle approximately one-third full. Remove the bottles from the pneumatic trough, and cover them with glass plates. Place them upright on the table until ready for use.

Describe the action that takes place when concentrated sulphuric acid is dropped into formic acid. Why is a water bath used?

Write the equation for the reaction.

What part does the sulphuric acid play in the reaction?

PROPERTIES OF CARBON MONOXIDE

1. Note the colour and odour (**CAUTION**) of the gas. Lower a lighted splint into the full bottle of the gas.

Describe the colour and odour of carbon monoxide.

Does the gas burn? Does it support combustion?

2. Pour 10 ml. of limewater into the second bottle, which contains a mixture of carbon monoxide and air, and shake. Bring a lighted splint to the mouth of the bottle. Observe the manner in which the gas burns, and shake the bottle again.

Has carbon monoxide any effect on limewater?

Describe the flame produced when a mixture of carbon monoxide and air burns.

What is the product of combustion of carbon monoxide and air? Why?

Write the equation for the burning of carbon monoxide.

QUESTIONS

1. Why is carbon monoxide so intensely poisonous if breathed continuously? (TEXT p. 274)

2. What are the properties that cause carbon monoxide to be such an insidious gas?

EXPERIMENT 71. To prepare and collect acetylene and study its properties.

PREPARATION AND COLLECTION

1. Half fill a 250 millilitre beaker with water, and into it, invert a test-tube filled with water. Drop a small lump of calcium carbide into the water in the beaker. Observe all changes that occur. Place the inverted test-tube filled with water over the calcium carbide, and collect the test-tube *full* of gas. Remove the test-tube and place it, mouth downward, on the table. Collect a second test-tube *half-full*, another *one-third full*, and a fourth approximately *one-twelfth full*, using additional lumps of calcium carbide if necessary.

2. Keeping the mouths of the test-tubes down, lift each one out of the beaker, allowing air to replace the water, and stand them mouth downward on the table. Save the contents of the beaker.

Describe the appearance of calcium carbide.

Make a list of all the observations in the above experiment.

Describe the colour and the odour of the gas produced when calcium carbide is dropped into water.

Write the equation for the laboratory preparation of acetylene.

PROPERTIES OF ACETYLENE

1. Hold each of the test-tubes, one at a time, in a nearly horizontal position, and bring a lighted splint to the mouth of each in turn. Observe closely the flame produced. After the reaction has taken place, add 5 ml. of limewater to each test-tube, and shake.

Describe and account for the way in which the acetylene burns in each case.

Why does the gas in the test-tube filled with acetylene burn so poorly?

Why is the action in the test-tube containing the least acetylene called complete combustion?

What evidence is there that complete combustion produces a higher temperature than does incomplete combustion? Explain.

What is the effect on the limewater in each case?

Write an equation to represent (a) the incomplete combustion, and (b) the complete combustion of acetylene.

2. Test the solution remaining in the beaker with a piece of neutral litmus paper. Filter some of the mixture, and collect about 5 ml. of the filtrate in a test-tube. Using a clean glass tube, blow your breath through the filtrate.

What does the litmus test indicate?

What is the filtrate? Why?

What does this experiment show regarding the solubility of calcium hydroxide in water?

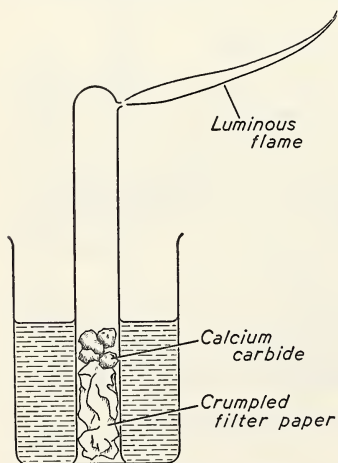


Fig. 61. An acetylene generator and burner.

3. Hold the ball of the thumb firmly against the mouth of a small test-tube (preferably soft glass), so as to seal the air within, and with a blow-pipe, direct the flame from a Bunsen burner against the rounded end of the test-tube. Heat a small area of the tube until the expanding air within, forces its way through the softened glass, leaving a small hole.

Place three or four small lumps of calcium carbide in the tube. Fold or roll a filter paper into a small ball, and plug the mouth of the test-tube with it *loosely*. Invert the test-tube in a beaker containing water to a depth of

approximately two inches (Fig. 61). After a minute or so, hold a lighted match to the small opening of the test-tube.

Explain why the flame is so regular in size. Why is the term "automatic" applicable to this burner?

QUESTIONS

1. Account for the difference in the construction of an acetylene burner used for lighting purposes, and one used for cutting and welding. (TEXT p. 276)

2. By means of suitable equations, show that decreasing the volume of oxygen in the combustion of acetylene results in an increase in the amount of carbon produced.

UNIT 12

Fuels and Flames

EXPERIMENT 72. The luminosity of flames.

1. Recall the burning of sulphur, and phosphorus, in air, and in oxygen (Experiment 13).

How does the flame temperature in pure oxygen compare with the temperature of the flame in air? Explain.

How is the luminosity of a flame affected by an increase in temperature?

2. Hold (a) a piece of copper wire, and (b) a piece of glass tubing, in the flame of a Bunsen burner, and observe the colour changes.

What is the effect of an increase in temperature on the luminosity of the copper and of the glass?

How is the luminosity of the flame affected by these substances? Explain.

3. Adjust the air inlet of your lighted Bunsen burner until a blue, non-luminous flame is produced. Support the burner in a horizontal position, and place a large asbestos pad underneath. Sprinkle each of the following substances into the flame, and observe the effect of each: (a) sodium nitrate, (b) potassium sulphate, (c) powdered calcium chloride, (d) powdered charcoal, (e) finely divided iron, and (f) zinc oxide.

Describe the effect produced on the burner flame by each substance.

In what respect is the result in (a), (b), and (c), different from the result in (d), (e), and (f)?

4. **Flame tests.** Because a non-luminous Bunsen flame produces a characteristic colour when some of the metals, such as sodium, potassium, calcium, barium, lithium, strontium, and copper (or their salts), are heated in it, *flame tests* are sometimes used as an aid in identifying these metals. Use soluble salts of each of the metals listed above.

Adjust the Bunsen burner so that a flame as nearly colourless as possible is produced. Such a flame will be a very faint blue. Hold a platinum wire which is embedded in a glass rod (Fig. 62) in the flame. If the flame becomes coloured, the wire is dirty. To clean the wire, dip it in hydrochloric acid, and heat it again. Continue

to dip the wire into the acid and reheat until the colour of the flame is not changed. Now dip the wire into a solution of the salt to be tested, and hold it in the flame. Note any colour change in the flame. Between each test, clean the wire by alternately dipping it into the acid, and heating it, until no colour is imparted to the flame.

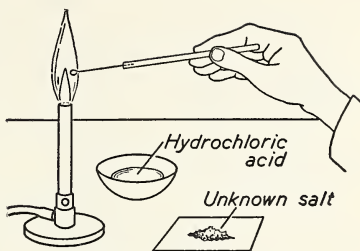


Fig. 62. Making a flame test.

What effect has the presence of a compound of (a) sodium, (b) potassium, (c) calcium, (d) barium, (e) lithium, (f) strontium, and (g) copper on a colourless flame?

5. Close the air inlet on the lighted burner. Observe the change in the luminosity of the flame. Hold a cold glass plate horizontally in the flame for a few seconds, and observe the surface of the glass.

Describe the appearance of the glass plate after it is removed from the flame. What is the deposit that forms on the plate? Account for its formation. Why does it form only when the air inlet is closed?

What substance causes the luminosity of the flame?

What happens to the solid material in the incomplete-combustion flame when the air inlet of the burner is opened?

EXPERIMENT 73. The products of the combustion of fuels.

1. Nearly fill a 250 millilitre flask with cold water. Make sure that the outside of the flask is wiped dry, and then hold it over the non-luminous flame of a Bunsen burner for a few seconds.

Describe the appearance of the outside of the flask.

2. Touch the deposit on the outside of the flask with a finger that has some anhydrous cupric sulphate adhering to it, or with a piece of dry cobalt chloride test-paper. (This paper may be made by dipping filter paper into a cobaltous chloride solution, and allowing it to dry.)

What is the liquid formed? What is the authority for your identification? Account for the formation of the liquid.

3. Over (a) a Bunsen flame, (b) a candle flame, and (c) burning kerosene in an evaporating dish, hold an inverted 250 millilitre beaker which has just been rinsed with limewater.

Describe and account for the change in appearance of the droplets of limewater.

QUESTIONS

1. Account for the production of heat when fuels burn.
2. What precautions should be taken to obtain the maximum heat value from a fuel?
3. Under what circumstances may a fuel be used for lighting purposes?
4. Name three fuels which are also used as illuminants.
5. Name the products resulting from the *complete* combustion of fuels that contain both carbon and hydrogen.
6. Name the products resulting from the *incomplete* combustion of fuels that contain both carbon and hydrogen.
7. Write equations for the *complete* combustion of each of the following fuels: coke, methyl alcohol (CH_3OH), gasoline (C_8H_{10}), water gas, and acetylene.

EXPERIMENT 74. The structure of flames.

1. Adjust the air supply on a Bunsen burner to produce a flame with a well defined inner cone. Hold a piece of cardboard *horizontally* in the flame, just below the tip of the inner cone. As soon as the cardboard begins to char, remove it from the flame, and examine it closely.

Hold another piece of cardboard *vertically* in the centre of the flame, resting it on the top of the barrel of the burner. When the cardboard begins to char, remove it from the flame, and examine it.

Make separate diagrams showing the areas charred by the flame in each case.

What do the charred areas indicate regarding the temperatures of the flame? Why is the flame hotter in some areas than in others?

2. Hold the end of a piece of glass tubing, approximately 6 inches long, inside the smaller inner cone of the Bunsen flame. Slope the tube, and bring a lighted match to the open end farthest from the flame.

What does the result of this experiment indicate regarding the nature of the inner cone? Is the temperature inside this inner cone high or low? Why?

3. Insert a pin through a "strike-anywhere" match just below the head, and place the match vertically in the barrel of an unlighted Bunsen burner, with the pin resting horizontally on the top of the barrel (Fig. 63B). Turn on the gas and light the burner.

Account for the behaviour of the match.

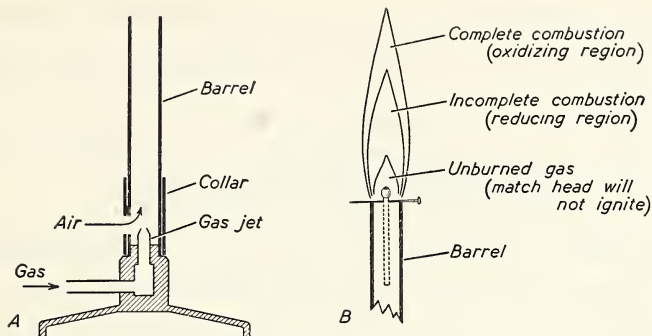


Fig. 63. A, Bunsen burner in section. B, diagram showing the various regions in a Bunsen flame.

4. Clean one end of a piece of copper wire 8 inches long, by dipping it into concentrated nitric acid and then washing it in water. Hold the clean end of the wire horizontally in the top portion of a Bunsen burner flame, and observe the effect on the copper. Lower the wire until it is in the central cone (reducing region), and observe the colour change in the copper. Remove the wire from the flame, let it cool in the air, and again observe the effect on the copper.

Describe and explain the colour change that takes place when the copper is heated in the tip of the flame.

Describe and explain the colour change that takes place when the copper is lowered into the central cone.

Describe and explain the action that takes place when the hot copper is removed from the flame and allowed to cool in air.

5. Light a Bunsen burner, and adjust the air supply to produce a well defined inner cone. Hold a piece of fine copper-wire gauze horizontally over the flame, and lower it to approximately one inch above the top of the barrel. Observe the effect on the flame.

6. Hold a piece of fine copper-wire gauze horizontally 1 inch above an unlighted Bunsen burner, turn on the gas supply, and bring a lighted match to the gas *above* the gauze. Observe the effect of the gauze on the flame.

Describe and account for the action of the gauze on the flame in each of these experiments.

Account for the use of the Davy safety lamp. Under what conditions would it cease to be effective?

7. Bring a lighted match down slowly to the wick of an unlighted candle, and observe closely the way in which the flame is produced.

When the flame of the candle is well established, blow it out, and immediately bring a lighted match to the wick.

Describe the action that takes place up to the time that the flame is well established.

What evidence is there that paraffin wax has a low melting point?

How does this property of paraffin make the use of a candle possible?

Why is the candle flame yellow in colour?

Why does the flame taper uniformly to a sharp point?

In what respects does a candle flame resemble the flame of a Bunsen burner?

Account for the action that takes place when the candle flame is re-lighted.

8. Make a helix of copper wire slightly larger in diameter and slightly longer than a lighted candle flame, and bring it over the flame.

Describe and account for the action of the flame.

9. Bring a helix of copper wire, which has been heated to redness in a Bunsen flame, over a lighted candle flame.

Describe and account for the action of the flame.

QUESTIONS

1. What factors are necessary for the production of a flame?
2. Why is the tip of a Bunsen burner flame an *oxidizing* flame?
3. Why is the middle cone (region of incomplete combustion) a *reducing* region?
4. What is the function of the barrel of a Bunsen burner?

UNIT 13

Sulphur and Its Compounds

EXPERIMENT 75. The allotropic modifications of sulphur and their properties.

1. Examine a small sample of (a) roll sulphur, and (b) flowers of sulphur.

Describe the colour and appearance of (a) roll sulphur, and (b) flowers of sulphur.

2. CAUTION: (*Demonstration*) Keep flames away. Carbon disulphide is volatile and very flammable.

Into a test-tube one-quarter full of carbon disulphide, place a piece of roll sulphur approximately the size of a bean. Cover the mouth of the test-tube with the thumb, and shake gently.

Is carbon disulphide a solvent for roll sulphur?

Continue to add either powdered roll sulphur or flowers of sulphur until no more will dissolve. Pour off the clear solution into a 50 millilitre beaker.

A few drops of this saturated solution may be placed on a microscope slide so that the crystals formed when the carbon disulphide evaporates may be viewed under the low power magnification of a microscope.

If larger crystals are desired, the beaker should be covered with several layers of filter paper held firmly in place by a glass plate. These crystals may be examined with a hand lens.

Describe the appearance of one of the more perfect crystals.

Name this allotropic modification of sulphur.

3. Half fill a Pyrex test-tube with small pieces of roll sulphur. Hold the test-tube at an angle of 45° , and heat the tube with a low Bunsen flame. Rotate the tube constantly and continue heating carefully until all the sulphur has just melted. Note the colour and nature of the liquid.

4. Pour the melted sulphur into a dry filter paper which has been placed in a funnel, and let it stand without jarring until there is evidence of crystals forming nearly to the centre of the sulphur surface.

Quickly pour the melted sulphur remaining in the filter paper into a beaker half full of cold water. Immediately remove the filter paper from the funnel, and open it to expose the crystals that have formed. Examine these crystals and keep them for Procedures 5 and 6.

Describe the changes that take place in the sulphur while it is being melted.

Which is denser, liquid sulphur or solid sulphur? Give a reason for your answer.

Does the melted sulphur wet glass?

Compare the viscosity of melted sulphur with that of water.

What is the nature of the crystals which form on the filter paper?

Name this allotropic modification of sulphur.

Describe the appearance of the sulphur formed when melted sulphur is poured into cold water.

Name this allotropic modification of sulphur.

5. CAUTION: Keep flames away. Carbon disulphide is volatile and very flammable.

Place a few of the crystals that formed on the filter paper (Procedure 4) in a test-tube. Add 2 ml. of carbon disulphide, and shake.

Describe the action of carbon disulphide on monoclinic sulphur.

6. Allow the remaining crystals of sulphur on the filter paper (Procedure 4) to stand until the next laboratory period, and observe any changes.

Describe and account for the changes observed.

7. Use the same test-tube as was used in Procedure 3, and half fill it

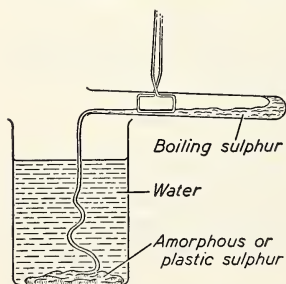


Fig. 64. Pouring boiling sulphur into cold water.

with roll sulphur. Heat the test-tube with a low Bunsen flame until the sulphur has melted. Continue heating carefully and slowly, tilting the tube occasionally to determine the viscosity of the sulphur, until the sulphur boils. Note the successive changes in colour and viscosity of the molten sulphur. Observe the inside of the test-tube immediately above the boiling sulphur. Pour the *boiling* sulphur slowly into a beaker full of cold water (Fig. 64). Examine the material in the beaker.

Describe the changes in colour and viscosity of the sulphur from the time that the solid first melts until the liquid sulphur boils.

Give two observations that would indicate that the sulphur is boiling.

Account for the appearance, colour, and formation of the sublimate.

What is its common name?

Describe the colour and nature of the sulphur resulting from pouring boiling sulphur into cold water. Name this allotropic modification of sulphur.

8. Divide the plastic sulphur, obtained by pouring boiling sulphur into cold water, into two portions. To one portion, in a test-tube, add 5 ml. of carbon disulphide (**CAUTION**), and shake. Filter, collect some of the filtrate on a watch-glass, and allow it to evaporate. Let the other portion of plastic sulphur stand until the next laboratory period, and examine it.

Is plastic sulphur soluble in (a) water, (b) carbon disulphide?

Describe and account for the change in appearance of the plastic sulphur when allowed to stand.

9. Place a small lump of roll sulphur, about the size of a pea, in an asbestos-lined deflagrating spoon, and ignite it. Observe the appearance of the flame and any changes in the sulphur itself. Cautiously note the odour of the gas (**Technique 14**). Lower the burning sulphur into a bottle of air containing 10 ml. of a dilute solution of potassium permanganate. Shake and observe.

Describe the appearance of the burning sulphur.

Describe the odour of the gas produced. Name this gas.

Describe and account for the change that takes place in the potassium permanganate solution.

Write chemical equations for (a) the burning of sulphur, and (b) the action of sulphur dioxide on potassium permanganate solution. (TEXT p. 307)

Why can the action with potassium permanganate be considered a chemical test for sulphur dioxide?

10. Fill the same test-tube that was used to prepare plastic sulphur one-quarter full of roll sulphur. Heat the sulphur until it boils. Keep it boiling, and lower either a spiral of fine copper wire or a strip of copper foil into the sulphur vapour. Observe the changes that take place.

Describe the nature and appearance of the copper before being lowered into the test-tube.

Describe and account for the reaction that takes place in the test-tube.

What is the nature and appearance of the product resulting from the reaction?

Compare the reaction between copper and sulphur with that between iron and sulphur (Experiment 44).

QUESTIONS

1. What physical property of sulphur makes possible its commercial recovery by the Frasch method? (TEXT p. 297)
2. Explain the statement "sulphur exists in three different allotropic modifications."
3. How could you prove that the allotropic modifications of sulphur have almost identical chemical properties?

EXPERIMENT 76. To prepare and collect hydrogen sulphide and study its properties.

PREPARATION AND COLLECTION

CAUTION: Unless exceptionally good ventilation equipment is available, it is inadvisable to perform this experiment in the laboratory, except by demonstration. It is best prepared in a fume cupboard.

Hydrogen sulphide is very poisonous. When inhaled in concentrations of one volume of the gas in two hundred volumes of air, it can be fatal. Bleaching powder sprinkled with acetic acid provides sufficient chlorine gas to serve as an antidote.

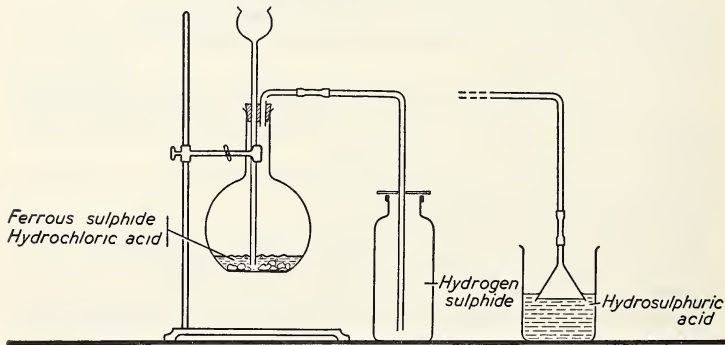


Fig. 65. The laboratory preparation and collection of hydrogen sulphide.

1. Arrange the apparatus as shown in Figure 65. Put several pieces of ferrous sulphide into the flask, and add dilute hydrochloric acid through the thistle-tube until the ferrous sulphide is well covered.
2. Collect five bottles full of the gas by the upward displacement of air. To determine when the bottles are filled with the gas, bring

to the mouth of each bottle a piece of filter paper which has been dipped into a lead acetate solution.

3. When the gas has been collected, prepare a solution of the gas in water as indicated in the diagram. Keep this solution for future use.

Describe the action that takes place within the flask.

State two reasons for collecting the gas by the upward displacement of air.

Write the equation for the laboratory preparation of hydrogen sulphide.

Describe the action that takes place when hydrogen sulphide reacts with lead acetate solution. Write the chemical equation to express this reaction. Explain why this reaction serves as a test for hydrogen sulphide.

PROPERTIES OF HYDROGEN SULPHIDE

1. Observe the colour and odour of the gas. Invert a bottle of the gas in a dish half full of water, and allow it to stand until the end of the period.

Describe the colour and the odour of hydrogen sulphide.

What does the height of the water in the bottle indicate about the solubility of hydrogen sulphide in water?

2. Into the second bottle of gas, drop (a) a piece of moist red litmus paper, (b) a piece of moist blue litmus paper, and (c) a silver coin.

Describe and account for the behaviour of each piece of litmus paper.

What change takes place in the silver coin?

What name is given to the solution of hydrogen sulphide in water?

3. Lower a blazing splint into the third bottle of hydrogen sulphide, and observe the reaction closely.

Does hydrogen sulphide burn?

Does hydrogen sulphide support combustion?

Describe the flame of the burning gas.

Describe the substance that forms inside the bottle. What is this substance? Account for its formation, and write the equation that represents this action.

4. Into the fourth bottle of the gas, pour 5 ml. of a solution of lead acetate.

Describe and account for the change in appearance of the lead acetate solution. Why does the solution in the bottle have a different appearance from the deposit on a piece of filter paper which has been dipped into a lead acetate solution and exposed to the gas?

Would the solution of any other soluble lead salt react in the same way?

5. Place a piece of sulphur, the size of a bean, in an asbestos-lined deflagrating spoon, ignite it, and lower it into a gas bottle filled with air. Let the sulphur burn until the flame is extinguished, and remove the deflagrating spoon from the bottle. Place this bottle of sulphur dioxide, mouth downward, over the remaining bottle of hydrogen sulphide, and let the bottles stand in this position for a few minutes. (If no reaction takes place, introduce a few drops of water to the mixture of gases.)

Describe and account for the action that takes place. Write the equation.

6. Test the solution of hydrogen sulphide in the beaker with a piece of neutral litmus paper. Pour some of the solution into a bottle, insert a stopper in the bottle, and allow it to stand until the next laboratory period.

Is the solution a strong or a weak acid? Explain.

Describe and account for the appearance of the hydrogen sulphide solution after standing for a day. Write an equation to represent the chemical change that has taken place.

7. CAUTION: Arsenic and antimony compounds are very poisonous.

In six separate test-tubes, place 10 millilitre samples of solutions of each of the following substances: cupric sulphate, lead nitrate, silver nitrate, zinc sulphate, arsenic trichloride, and antimony trichloride. Add in turn to each test-tube, 2 ml. of freshly prepared hydrogen sulphide solution, and observe the changes that take place. Observe the effect of adding 2 ml. of ammonium hydroxide to the test-tube that contained the zinc sulphate solution.

Describe and account for the changes that take place in each test-tube when hydrogen sulphide solution is added. Write a chemical equation to represent each reaction.

What is the effect of adding ammonium hydroxide to the test-tube containing the zinc sulphide? In what way does zinc sulphide differ from the other sulphides produced in this experiment?

QUESTIONS

1. Why is the burning of a full bottle of hydrogen sulphide an example of incomplete combustion?

2. Write the chemical equation to represent (a) the *complete* combustion of hydrogen sulphide, and (b) the *incomplete* combustion of hydrogen sulphide.

3. Some scientists believe that the sulphur deposits of Louisiana

have been formed by the interaction of hydrogen sulphide gas and sulphur dioxide gas. If this did take place, account for the formation of the hydrogen sulphide and the sulphur dioxide required for these vast deposits. Write a chemical equation that would indicate the reaction that took place in the formation of the sulphur.

EXPERIMENT 77. To prepare and collect sulphur dioxide and study its properties.

PREPARATION AND COLLECTION

CAUTION: Unless good ventilation equipment is available, it is inadvisable to perform this experiment in the laboratory, except by demonstration.

1. Arrange the apparatus as shown in Figure 66, and place a concentrated solution of sodium bisulphite in the flask. Control the rate at which sulphuric acid drops into the flask, so that a steady flow of gas is produced.

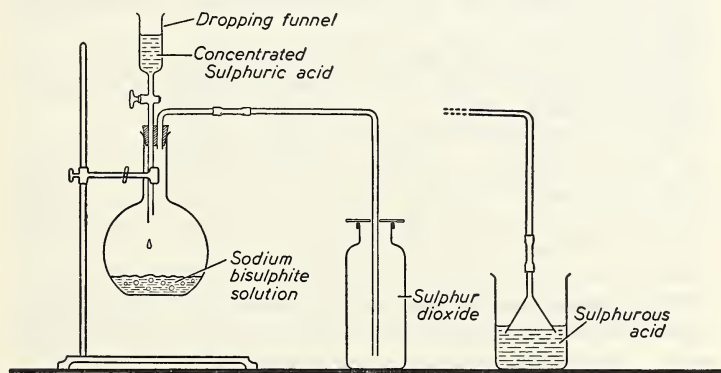


Fig. 66. The laboratory preparation and collection of sulphur dioxide.

2. Collect four bottles of the gas by the upward displacement of air, and place a glass plate over the mouth of each bottle.

3. Place a funnel on the end of the delivery tube as illustrated in the diagram, and make a solution of sulphur dioxide in water.

Describe the action that takes place when the acid and the sodium bisulphite solution react.

Write the chemical equation for the laboratory preparation of sulphur dioxide.

PROPERTIES OF SULPHUR DIOXIDE

1. Observe the colour and odour (*Technique 14*) of the gas. Note the effect on moist neutral litmus paper.

Describe the colour and odour of sulphur dioxide.

What change takes place in the neutral litmus paper? Account for the action of sulphur dioxide on litmus.

2. Lower a blazing wooden splint into a bottle of sulphur dioxide.

Does the gas burn? Does the gas support combustion? Explain how advantage may be taken of these properties to extinguish a fire in a chimney.

3. Pour 5 ml. of a dilute solution of potassium permanganate into the third bottle of sulphur dioxide. Add 2 ml. of barium chloride solution, and observe the effect produced. Finally, add 2 ml. of dilute hydrochloric acid, and observe the effect.

Describe and account for the change in appearance of the potassium permanganate solution. Why does this reaction serve as a chemical test for sulphur dioxide?

Write the equation for the action of potassium permanganate on sulphurous acid.

Describe the action of the barium chloride solution.

Write the equation for the action of barium chloride on the solution remaining in the gas bottle.

Why was hydrochloric acid added? (TEXT p. 312)

4. Place a thoroughly moistened, coloured flower (e.g., a red rose) in the fourth bottle of sulphur dioxide. After a short time, observe the colour of the flower, and then hold it in the vapour of concentrated nitric acid. Observe any further change.

What is the effect of sulphur dioxide on the moist flower? Explain.

What happened when the flower was held in nitric acid fumes? Explain.

Substances bleached with sulphur dioxide often return to their original colour. Explain with the aid of a suitable equation.

5. Pour 5 ml. of the sulphurous acid solution contained in the beaker into a test-tube, and add a small piece of magnesium ribbon.

Describe and account for the action that takes place.

How would you identify the gas that is given off?

What product remains in the solution?

Write the chemical equation for the reaction.

6. Pour 5 ml. of the sulphurous acid solution into a test-tube, add 5 ml. of a solution of barium chloride, and shake. To the contents of the test-tube, add dilute hydrochloric acid, a few drops at a time, until no further change takes place.

Describe and account for the action of the barium chloride solution with sulphurous acid. Write the equation for the reaction.

Describe and account for the action when hydrochloric acid is added to the contents of the test-tube. Write the equation for the reaction.

7. Allow the remainder of the solution to stand in a beaker exposed to the air until the next laboratory class, and repeat the test with barium chloride solution and hydrochloric acid, as in Procedure 6.

Describe and account for the action of the barium chloride solution. Explain with the aid of equations.

Account for the action with hydrochloric acid.

How could you distinguish a sulphite from a sulphate?

EXPERIMENT 78. (*Demonstration*) A method of preparing sulphuric acid in the laboratory.

1. Select a three-hole rubber stopper to fit an 8 oz. gas-collecting bottle. Insert a short piece of glass tubing, which will act as an outlet tube, into one of the openings. Insert a glass delivery tube leading from a sulphur dioxide generator (Experiment 77) into a second opening. Insert a glass delivery tube, leading from a Pyrex test-tube containing four grams of lead nitrate, into the third opening.

2. Heat the lead nitrate until the gas bottle is filled with brown fumes; then allow the sulphuric acid to drop slowly into the flask containing the sodium bisulphite solution. Control the heating of the lead nitrate so as always to have some brown gas in the gas bottle.

3. After a short time, observe and describe the appearance of the inside of the gas bottle.

4. When brown fumes are no longer produced in the test-tube, disconnect the apparatus, and slowly, a few drops at a time, add 10 ml. of water to the gas bottle.

Describe the action that takes place when water is added to the gas bottle.

5. Divide the solution produced in Procedure 4 into two parts, by pouring it into two test-tubes. Add a solution of barium chloride, followed by 2 ml. of hydrochloric acid, to one test-tube. Add a small piece of mossy zinc to the other test-tube, and test with a blazing splint.

Describe and account for the action with the barium chloride solution and the hydrochloric acid.

Describe and account for the action with the zinc.

QUESTIONS

1. Name the brown gas produced when lead nitrate is heated. Write the equation for the reaction.
2. What is the function of the brown fumes in the reaction? Would the oxygen produced by the heating of lead nitrate help or hinder the reaction? Explain.
3. Write the equation for the production of sulphur trioxide.
4. Describe sulphur trioxide. What happens when water is added to the sulphur trioxide in the gas bottle? Write the equation.

EXPERIMENT 79. The properties of sulphuric acid.

CAUTION: Concentrated sulphuric acid is very destructive to flesh and clothing and should be handled with care.

1. (*Demonstration*) Put 40 ml. of water into a 250 millilitre beaker. Place in the water a thermometer calibrated to read up to $200^{\circ}\text{C}.$, and record the temperature. With constant stirring, pour 100 ml. of concentrated sulphuric acid into the water, in a fine stream. Record the temperature.

Describe the action that takes place when concentrated sulphuric acid is added to water.

What temperature change is indicated by the thermometer?

Explain why water must never be added to concentrated sulphuric acid.

Why is concentrated sulphuric acid a good drying agent?

2. (*Demonstration*) Half fill a 125 millilitre beaker with cane sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$). Add just sufficient concentrated sulphuric acid to cover the sugar, and observe. (**CAUTION:** The spongy carbon mass left at the end of this experiment often contains some trapped concentrated acid. Pupils should not be permitted to handle it.)

Describe the changes that take place, and account for these changes.

Name the products resulting from the reaction.

Write a chemical equation to represent the reaction. What role does sulphuric acid play in the reaction?

3. Dip a wooden splint into concentrated sulphuric acid and observe the effect. Wood contains cellulose ($\text{C}_6\text{H}_{10}\text{O}_5$)_n.

Describe and account for the change that takes place.

Write a chemical equation to represent the reaction that takes place.

4. Dip a glass tube into concentrated sulphuric acid, and trace a few lines with the concentrated acid on a piece of filter paper.

Describe and account for the action of the concentrated acid on the paper.

Write an equation to represent the reaction.

5. Dip a glass tube into dilute sulphuric acid, and trace a few lines with the dilute acid on a piece of filter paper. Warm the filter paper over a Bunsen flame, and observe the action that takes place.

What is the action of the dilute sulphuric acid on the filter paper? Why is the filter paper warmed?

6. (CAUTION) Place two drops of concentrated sulphuric acid on a piece of mica, and heat the mica with a Bunsen flame.

Describe and account for the changes that take place.

7. (Demonstration) Into each of two separate test-tubes, place a piece of mossy zinc. Add 5 ml. of dilute sulphuric acid to one test-tube, and 5 ml. of concentrated sulphuric acid to the other. Warm the test-tube containing the concentrated acid.

Describe and account for the gaseous product formed in each case.

Write an equation for each reaction.

UNIT 14

Sodium Chloride and Hydrogen Chloride

EXPERIMENT 80. The solubility and crystal structure of sodium chloride.

1. Select two test-tubes of the same size. Half fill one of them with cold water, and half fill the other with hot water. To each test-tube add one level teaspoonful of sodium chloride. Shake the tubes vigorously, and allow them to stand until any undissolved salt settles.

What kind of solution is present in each test-tube?

Is there a noticeable difference in the amount of sodium chloride that remains undissolved in each case? Describe and explain.

What is the solubility of sodium chloride in water at (a) 10°C., (b) 90°C.? (TEXT p. 97)

2. Filter both solutions through a common filter paper. Collect the filtrate, and divide it into two parts. By means of a Bunsen flame, rapidly evaporate one portion to dryness. Set the other portion aside and let it evaporate to dryness slowly. Examine the contents of each container. Select one large crystal and examine it with a magnifying glass.

What is the appearance of the residue resulting from (a) rapid evaporation, (b) slow evaporation? Account for the difference.

Describe the crystal structure of sodium chloride.

QUESTIONS

1. List four properties of sodium chloride.
2. What property of sodium chloride makes its commercial recovery possible?
3. Suggest a method of obtaining a large crystal of sodium chloride.

EXPERIMENT 81. (*Demonstration*) The electrolysis of an aqueous solution of sodium chloride (brine).

1. Set up the Hoffman electrolysis apparatus as shown in Figure 42, and fill it with a saturated solution of sodium chloride. Pass a direct current of electricity through the solution (see Experiment 35),

Observe closely the action that takes place at each electrode. When sufficient gas has collected in each arm of the apparatus, place a test-tube over each jet, and collect a test-tube full of each gas.

Describe the action at the cathode.

Describe the action at the anode.

What is the colour of each gas collected?

2. Bring a lighted splint to the test-tube filled with the gas collected over the cathode.

What is the name of this gas?

3. Note carefully the odour of the gas collected over the anode, and test the gas with a piece of moist blue litmus paper.

Describe the odour of the gas.

Describe the effect of the gas on moist blue litmus paper.

What is the name of this gas?

4. Place a few drops of the solution remaining in the electrolysis apparatus on a piece of red litmus paper. Test a solution of sodium chloride with a piece of red litmus paper.

Account for the difference in the litmus tests.

What new chemical is present in the solution remaining in the electrolysis apparatus? Account for its formation.

QUESTIONS

1. Name three important substances resulting from the electrolysis of sodium chloride solution. Explain the changes that take place to produce these substances. Write a summation equation representing the reactions that take place.

2. Name four commercial uses for each of the products formed by the electrolysis of brine. (TEXT p. 320)

EXPERIMENT 82. To prepare and collect hydrogen chloride and study its properties.

PREPARATION AND COLLECTION

1. Set the apparatus as shown in Figure 67. Put approximately four teaspoonfuls of sodium chloride into the flask and add just sufficient water to moisten it. Through the thistle-tube, add enough concentrated sulphuric acid to cover the end of the thistle-tube.

2. Heat gently. When necessary to the maintenance of the production of gas, add more acid.

3. Collect two bottles and two test-tubes of the gas by the upward displacement of air. Then, to obtain a solution of hydrogen chloride in water, arrange the funnel as shown in the diagram.

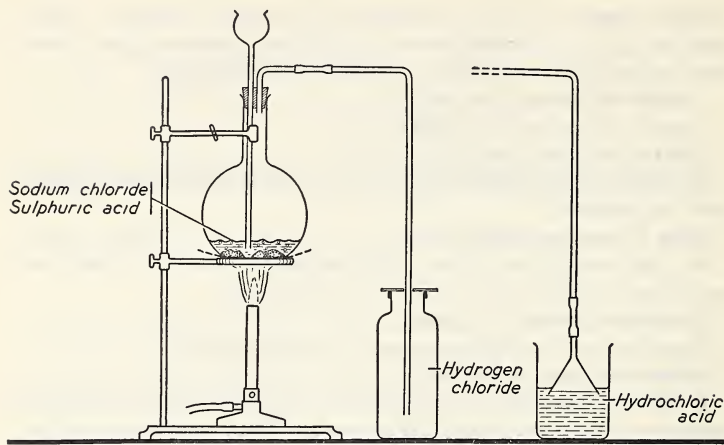


Fig. 67. The laboratory preparation and collection of hydrogen chloride.

Describe the action in the flask when the acid is added. What is the effect of heating the mixture?

What is the colour of the gas?

What does the method of collection indicate regarding the density of the gas?

Write the equation for the laboratory preparation of hydrogen chloride.

PROPERTIES OF HYDROGEN CHLORIDE

1. Invert a test-tube of the gas in a beaker of water.

Describe the action that takes place.

What conclusion can be made regarding the solubility of hydrogen chloride in water?

2. Note the odour of the gas in the second test-tube by wafting small quantities towards the nose. Gently breathe out across the mouth of the tube.

Describe the odour of the gas.

Describe and explain the action that takes place when the moist breath comes into contact with hydrogen chloride.

3. Insert a blazing splint into one of the bottles of gas.

Does the gas burn?

Does the gas support combustion?

4. Moisten a piece of filter paper with ammonium hydroxide, and drop the paper into the remaining bottle of gas.

Describe and account for the action that takes place.

5. Add a piece of neutral litmus paper to the solution of hydrogen chloride in water. Fill four test-tubes one-quarter full of the solution, and test in turn with each of the following substances: magnesium, zinc, washing soda crystals, and silver nitrate solution. In the experiments with magnesium and zinc, test the gas produced with a blazing splint.

What is the effect of the solution of the gas on litmus? Name this solution.

Describe the action of magnesium and zinc on the solution. Which metal seems more active? What gas is produced? Write equations for the reactions.

Describe and account for the action that takes place with washing soda crystals. Write the equation for the reaction.

Describe and account for the action with silver nitrate solution. Write the equation for the reaction.

EXPERIMENT 83. (*Demonstration*) A hydrogen chloride fountain.

Make a jet at one end of a piece of glass tubing about $1\frac{1}{2}$ feet long. Insert this tubing through one opening in a two-hole stopper, and a medicine dropper filled with water through the other opening. Allow the untapered end of the glass tubing to extend nearly to the bottom of a large beaker filled with a blue litmus solution.

Select a sturdy, round-bottomed flask and fill it with hydrogen chloride gas by the upward displacement of air. When the flask is filled, quickly insert the stopper into the inverted mouth, and support the flask in a retort ring as shown in Figure 68. Force a few drops of water out of the medicine dropper into the flask, and observe the action that takes place.

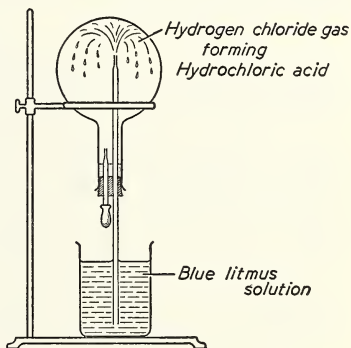


Fig. 68. A hydrogen chloride fountain.

Explain the formation and colour of the fountain.

QUESTIONS

1. List six properties of hydrogen chloride gas.

2. Give two identifying tests for hydrogen chloride, and write chemical equations for these tests.

3. Can "fogging of the breath" be used as an identifying test for hydrogen chloride? Explain.

UNIT 15

The Halogens

EXPERIMENT 84. Common methods for the preparation of chlorine.

CAUTION: Chlorine gas is a lung-irritant and because of its poisonous nature should never intentionally be inhaled.

1. Place about 1 gm. of sodium chloride and 1 gm. of manganese dioxide in a test-tube, and shake to mix the chemicals thoroughly. Add a few drops of concentrated sulphuric acid to the mixture. Note the colour of the gas produced and smell it *cautiously*. Test the gas with *moist* red and blue litmus papers. Blow your breath gently across the mouth of the test-tube.

Describe the action that takes place.

What is (a) the colour, (b) the odour of the gas?

What changes take place in the litmus papers? Has any gas prepared previously had this same effect on litmus?

Describe the action that takes place when the breath is blown across the mouth of the test-tube. What gas produces this effect? Is pure chlorine produced by this method, or is it mixed with another gas? Explain.

2. Place about 1 gm. of manganese dioxide in a clean, dry test-tube, and add enough concentrated hydrochloric acid to cover it. Warm the mixture, and test the gas as in Procedure 1.

Is the gas coloured? Is it mixed with the same gas as in Procedure 1?

3. Place a few crystals of potassium permanganate in a test-tube, and add a few drops of concentrated hydrochloric acid.

Identify the gas formed in this reaction.

4. Place about 2 gm. of bleaching powder in a test-tube, add sufficient dilute sulphuric acid to cover the powder, and warm the mixture gently.

What gas is produced? How did you identify it?

QUESTIONS

1. Which of the above methods is most suitable for the laboratory preparation of chlorine? Why?

2. What is the chief objection to using Procedure 1 as a laboratory method for preparing chlorine?
3. Write chemical equations to represent each of the four methods used to prepare chlorine.
4. List three properties of chlorine as determined by this experiment.

EXPERIMENT 85. (*Demonstration*) To prepare and collect chlorine and study its properties.

PREPARATION AND COLLECTION

1. Set up the apparatus as shown in Figure 69. Put two teaspoonfuls of manganese dioxide into the flask, and pour sufficient concentrated hydrochloric acid down the thistle-tube to cover the solid completely. Warm the contents of the flask, and collect the gas in gas bottles by the upward displacement of air, as shown in the figure.

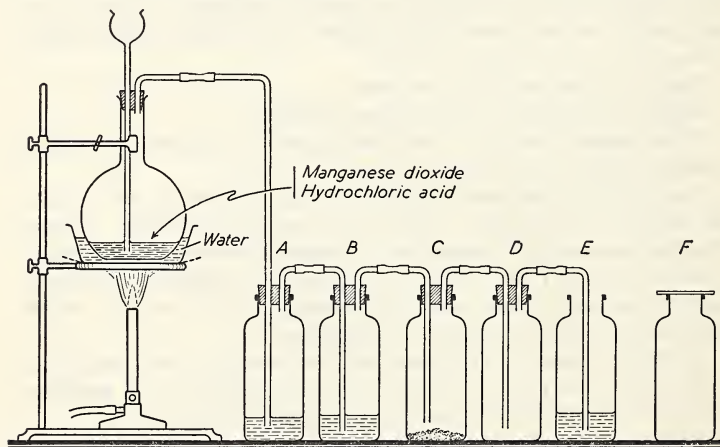


Fig. 69. The laboratory preparation and collection of chlorine. *A*, contains water to prepare chlorine water and to dissolve any hydrogen chloride that may be driven over. *B*, contains concentrated sulphuric acid to dry the chlorine. *C*, contains slaked lime for making a sample of bleaching powder. *D*, is for the collection of pure dry chlorine. *E*, contains a solution of sodium thiosulphate to prevent the escape of chlorine gas into the laboratory. *F*, a bottle of dry chlorine that has been filled at *D*.

2. Collect eight small bottles, and one extra large bottle to be used in Procedure 6. Cover each bottle with a glass plate, and allow them to stand on the desk until needed.

What does the method of collection indicate regarding the density of chlorine?

What observations are noted in the flask itself?

What method is used for determining when the bottles are filled with the gas?

PROPERTIES OF CHLORINE

1. Carefully dry a small piece of sodium with filter paper to remove all traces of adhering liquid, and then flatten the sodium into a thin wafer. Place this thin wafer in a deflagrating spoon, and lower it into a bottle of chlorine gas. **(If great care is used, the product of this reaction may be tasted.)**

Describe the action that takes place.

Describe and identify the product of the reaction.

Why must great care be used in tasting the product?

Write the chemical equation for the reaction.

2. Cut a strip of copper foil one-half inch wide and six inches long, and after heating it in the flame of a Bunsen burner, drop it quickly into a bottle of chlorine gas. After the reaction ceases, add 5 ml. of water to the bottle.

Describe the action that takes place.

Describe and name the product of the reaction.

What evidence is there that copper is less active chemically than sodium?

Describe and account for the colour change when water is added.

Write the chemical equation for the action of copper on chlorine.

3. Place a small piece of white phosphorus, about the size of a match-head, in an asbestos-lined deflagrating spoon, and lower it into a bottle of chlorine gas. After the reaction has ceased, add 5 ml. of water to the bottle, and shake.

Describe the action that takes place.

Describe and name the products of the reaction.

Are the products of the reaction soluble in water? Explain.

Could this reaction correctly be classified as combustion? Explain.

What does this reaction indicate regarding the properties of (a) phosphorus, (b) chlorine?

Write the equations for the reaction.

4. Place a pinch of powdered antimony in the fold of a small piece of paper, and slowly sprinkle it into a bottle of chlorine.

Describe the action that takes place.

Describe and name the products of the reaction.

What evidence is there that the reaction is exothermic?

Write the equations for the reaction.

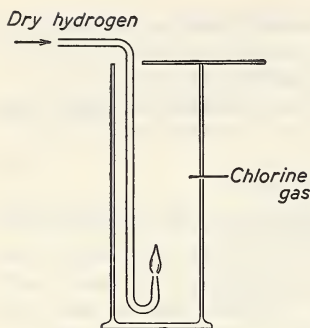


Fig. 70. The combustion of hydrogen in chlorine.

5. Set up a hydrogen generator, and attach a glass jet of the shape shown in Figure 70. When an inverted test-tube filled with hydrogen from this jet burns quietly, indicating the evolution of pure hydrogen, light the jet with the burning hydrogen, and lower it into a bottle filled with chlorine. When the flame goes out, remove the jet, and fill the hydrogen generator with water. Insert a piece of moist neutral litmus paper into the bottle, and blow your breath

across its mouth.

Describe the change in the appearance of the hydrogen flame when burning in chlorine.

Account for the change in colour noted in the gas bottle.

What causes the flame to be extinguished? What gas is now present in the gas bottle? Explain.

Write the chemical equation for the reaction.

6. Dip a piece of filter paper into warm turpentine, wind it into a loose roll sufficiently small to fit the mouth of the **large** bottle of chlorine, and quickly insert it into the bottle. After the reaction is over, blow gently across the mouth of the bottle.

Why is it necessary to use a large bottle full of chlorine?

Describe the action that takes place. Is it exothermic?

What gas must now be present in the bottle? Explain.

Name and account for the formation of the other product of this combustion.

What important property of chlorine is demonstrated here?

Write the chemical equation to represent this reaction.

7. Lower a burning candle or a wax taper into a bottle of chlorine, and note the effect on the appearance of the flame.

Describe the change in the appearance of the flame.

Account for the reaction that takes place.

What important property of chlorine is demonstrated by this experiment?

8. Into another bottle of chlorine, place a piece of lined foolscap upon which has been scribbled some pencil and ink marks. Also

insert red and blue litmus papers, a cancelled postage stamp, and a piece of fruit-basket netting or other similarly dyed cloth.

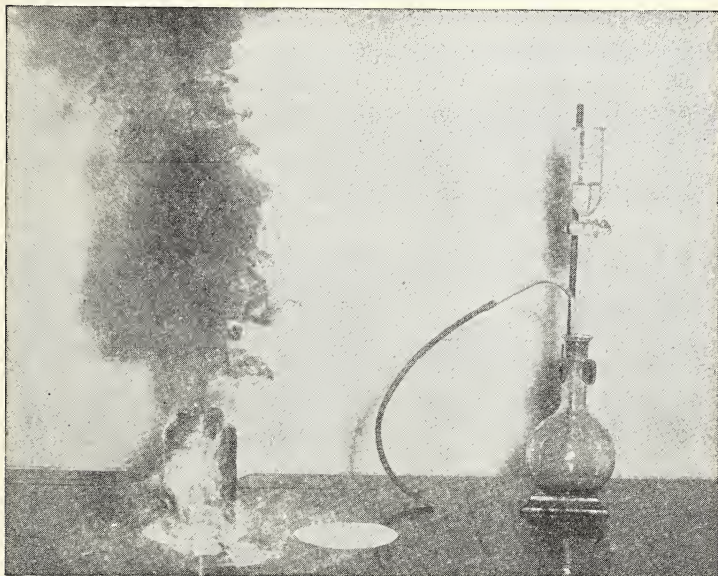
Describe the effect of the dry chlorine on each of these materials.

9. Repeat Procedure 8, this time dipping each material into water before inserting it into the chlorine bottle.

Describe the effect of the chlorine on the moistened materials.

Account for any difference in action when water is present.

Write a summation equation to explain this important chemical property of chlorine. (TEXT p. 330)



Paul Kyselka

Fig. 71. The combustion of turpentine in chlorine. When a filter paper dipped in warm turpentine is placed in a bottle of chlorine gas, the turpentine burns with an orange-coloured flame, producing a dense cloud of carbon particles. The hydrogen chloride gas which is also produced causes the white fumes around the gas bottle.

10. Place a piece of starch-iodized paper into a bottle of chlorine and observe the colour change in the paper. This colour change (especially when coupled with chlorine's colour, odour, and bleaching properties) may be considered a characteristic test for the gas.

What is the effect of chlorine on starch-iodized paper?

11. Remove the gas bottle into which slaked lime was placed (Fig. 69C), and examine the product resulting from its contact with

chlorine. Place half of the material in a test-tube, and add a small amount of dilute sulphuric acid. Note the odour of any gas liberated, and test it with moist blue litmus paper.

Make a thin paste by adding water to the remaining portion of the material. Dip a piece of fruit-basket netting into the paste, and then drop it into dilute sulphuric acid in a beaker. Remove the cloth from the beaker, wash it in water, and observe its colour.

What is the product formed by passing chlorine over slaked lime?

Why is this product an important industrial compound? How is chlorine liberated from this compound?

Write equations to represent (a) the formation of bleaching powder, and (b) the liberation of chlorine from bleaching powder.

12. To the liquid in bottle A (Fig. 69), add blue and red litmus papers and a piece of coloured cotton cloth. Set the bottle aside for five minutes, and examine.

What is the colour of the solution of chlorine and water?

What colour changes occur in the litmus and the cloth?

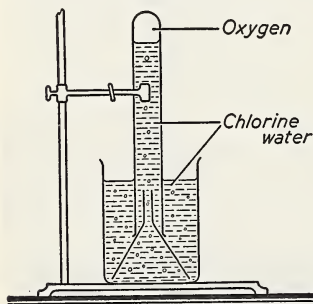


Fig. 72. The action of sunlight on a solution of chlorine and water.

13. Remove the chlorine generator from the system of gas bottles, and bubble chlorine through water contained in a 400 millilitre cylinder until the solution is saturated with chlorine. Fill a large test-tube with this solution, and invert it over a funnel, as shown in Figure 72. Allow the apparatus to stand in sunlight, and after all action ceases, test the gas which collects in the test-tube with a glowing splint. Test the liquid remaining in the beaker with blue litmus paper.

Describe the action that takes place when sunlight falls on a freshly prepared solution of chlorine and water.

What is indicated by (a) the glowing splint test (b) the litmus test? Account for the formation of each substance.

Write equations to represent the above chemical reactions.

QUESTIONS

1. Make a diagram showing how to set up the apparatus for obtaining chlorine by the action of potassium permanganate on hydrochloric acid, and write the chemical equation for the reaction.

2. List five physical properties of chlorine.
3. List four chemical properties of chlorine.
4. What is the most important chemical property of chlorine?
5. How does bleaching with chlorine differ from bleaching with sulphur dioxide? Explain with suitable equations.
6. Explain the statement "chlorine is made by the oxidation of hydrochloric acid using manganese dioxide as the oxidizing agent."
7. What important property of chlorine is indicated by the starch-iodized paper test? Explain with the aid of a suitable equation.

EXPERIMENT 86. To prepare and collect bromine and study its properties.

CAUTION: The vapour of bromine must not be inhaled, as it is a lung-irritant. Liquid bromine causes serious flesh burns.

PREPARATION AND COLLECTION

1. Arrange the apparatus as in Figure 73. Mix thoroughly on a piece of paper one teaspoonful of sodium bromide and one teaspoonful of manganese dioxide. Place the mixture in a large test-tube, and add just sufficient water to moisten it. Add just sufficient concentrated sulphuric acid to cover the moist mixture. Quickly insert the stopper containing the delivery tube leading to the gas bottle, and observe the reaction. When the reaction diminishes in activity gently warm the test-tube.

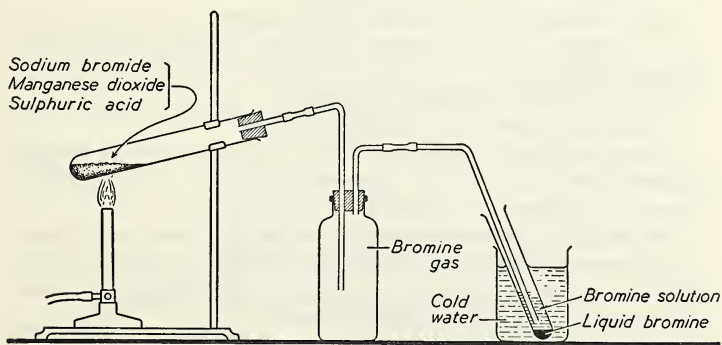


Fig. 73. The laboratory preparation and collection of bromine.

2. As the bottles become filled, remove them. Collect four bottles of bromine vapour. Place a glass plate on each bottle, and reserve until ready for use.

What advantage is there in moistening the solids with water before the acid is added?

Describe the action that takes place within the test-tube.

Is there any evidence that hydrogen bromide is produced? Explain.

Suggest a reason for collecting the bromine as indicated in Figure 73.

Write the chemical equation for the laboratory preparation of bromine.

PROPERTIES OF BROMINE

1. Examine the bromine in (a) the gas bottle, and (b) the test-tube.

What is the colour of bromine vapour?

What is the colour of liquid bromine?

What is the colour of bromine water?

Why is the gaseous state of bromine referred to as bromine vapour?

(TEXT p. 9)

Describe the odour of the vapour.

2. Remove the glass plate from a bottle of bromine vapour. Place moist blue and red litmus papers in the bottle, and allow it to stand.

Describe and explain the action on the litmus papers.

NOTE: For the following experiments, it may be more convenient to use a laboratory supply of bromine.

3. To a bottle of bromine vapour add water to a depth of one inch, and shake vigorously. Drop a piece of blue litmus paper and a piece of red litmus paper into the solution.

What evidence is there to show that bromine is soluble in water?

Explain the changes in colour of the litmus papers.

4. To a bottle of bromine vapour, add 5 ml. of carbon tetrachloride, and shake vigorously.

Is bromine soluble in carbon tetrachloride?

How does the action compare with that of water?

5. Sprinkle a pinch of powdered antimony into a bottle of bromine vapour.

Describe the action that takes place.

How does this action compare with the action of antimony on chlorine?

6. (**Demonstration**) Place a small piece of white phosphorus, about the size of a match-head, in an asbestos-lined deflagrating spoon, and lower it into a bottle of bromine vapour.

Describe the action that takes place.

What are the nature and the names of the products?

7. Carefully remove the test-tube containing the liquid bromine, from the beaker. Note the colour of the bromine solution above the liquid bromine. Add a few drops of bromine to a small portion of starch paste.

Why is caution necessary in removing liquid bromine from a container?

Describe the appearance of the starch paste before and after the addition of bromine.

8. Put 3 ml. of bromine water into a test-tube, and drop a piece of starch-iodized paper into the solution.

Describe and account for the colour change in the starch-iodized paper.

Can this colour change be used as a distinctive test for bromine, as it is for chlorine? Explain.

What does this action of bromine indicate regarding the relative activities of bromine and iodine? Explain with the aid of a chemical equation.

9. Into a test-tube half-filled with water, place a level teaspoonful of sodium bromide, and shake. To the resulting solution, add 2 ml. of carbon tetrachloride. Then add freshly prepared chlorine water to fill the test-tube three-quarters full, and shake vigorously.

What is the colour of sodium bromide solution?

What is the colour of carbon tetrachloride after it is placed in the sodium bromide solution?

What happens when the chlorine water is added? Explain.

What happens when the test-tube is shaken? Explain.

What does this experiment indicate about the relative activities of chlorine and bromine? Explain with the aid of a chemical equation.

QUESTIONS

1. List the physical and chemical properties of bromine.
2. What experimental evidence is there to indicate that bromine is less active chemically than chlorine?

EXPERIMENT 87. To prepare and collect iodine and study its properties.

PREPARATION AND COLLECTION

Place a teaspoonful of potassium iodide and a teaspoonful of manganese dioxide in a clean, dry, 250 millilitre beaker, and shake to mix the two chemicals. Fill an evaporating dish almost full of cold water, and arrange the beaker and evaporating dish as in Figure 74.

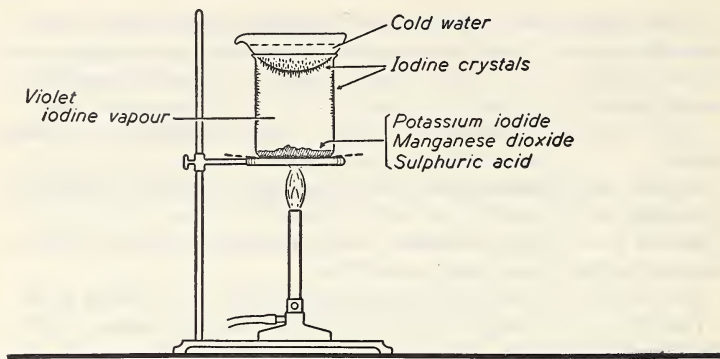


Fig. 74. The laboratory preparation and collection of iodine.

Lift the evaporating dish, and add approximately 10 ml. of concentrated sulphuric acid to the mixture in the beaker. Carefully replace the evaporating dish, and apply very gentle heat to the beaker for a short time. Then remove the flame.

Describe the action that takes place inside the beaker.

What is the colour of iodine vapour?

What is the colour of iodine crystals? Where are the crystals found? Account for their formation.

Write the chemical equation for the laboratory preparation of iodine.

PROPERTIES OF IODINE

1. Place a few crystals of iodine in a clean, dry test-tube. Hold the test-tube at a 45° angle and heat gently.

Describe all the changes that take place in the iodine, and account for these changes.

Why is the iodine heated gently? Why is the test-tube held at an angle?

Describe the crystal structure of iodine.

2. Place a few crystals of iodine in each of seven clean, dry test-tubes. Then, using one test-tube at a time, test the solubility of iodine in each of the following liquids: water, a solution of potassium iodide, ethyl alcohol, ether, chloroform, carbon tetrachloride, and carbon disulphide. Retain each of the solutions for Procedure 5.

Describe the visible changes in each case.

Which solvent seems to be (a) the best, (b) the poorest?

Could solubility in any of these solvents be used as an identifying test for iodine? Explain.

3. Place a few crystals of iodine in a mortar together with a small drop of mercury, and grind the two together with a pestle.

CAUTION: The product of this reaction is a poison.

Describe and interpret the result.

Write the chemical equation for the reaction.

4. (**Demonstration**) Place a piece of white phosphorus, about the size of a grain of wheat, in a clean, dry evaporating dish, and carefully sprinkle a few crystals of iodine on the phosphorus.

Describe and interpret the result.

Write the chemical equation for the reaction.

5. Put a teaspoonful of corn starch into a 250 millilitre flask, and half fill the flask with water. Heat the starch-water mixture until it appears translucent. Add sufficient water to form a mixture that pours easily. Test separate quantities of this starch paste or "solution" with each of the seven solutions retained from Procedure 2. If necessary, dilute the contents of each test-tube with water until a recognizable colour is produced.

Account for the change in appearance of the starch when heated with water.

Describe the colour effect in each test-tube.

What is the effect of water dilution on the colour?

Account for the starch-iodized paper test used with chlorine and bromine.

How would you test for the presence of starch in a food sample?

QUESTIONS

1. What procedure would be followed in purifying a sample of iodine?

2. How would you distinguish between separate solutions containing chlorine, bromine, and iodine?

3. Compare the properties of chlorine, bromine, and iodine, showing how the properties change with increasing atomic numbers. (TEXT p. 343)

EXPERIMENT 88. Tests for chlorides, bromides, and iodides.

1. Into each of three separate test-tubes, put (a) sodium chloride, (b) sodium bromide, and (c) sodium iodide, to a depth of one-quarter inch, and to each test-tube, add sufficient distilled water to fill it three-quarters full. Divide each solution into three parts, and arrange three sets of solutions so that each set contains solutions of the three salts. Use one set of solutions in each of the following procedures.

2. Bubble chlorine (or add freshly prepared chlorine water) into each solution in turn. After any visible changes have taken place, add a small amount of carbon tetrachloride to each test-tube.

Describe the visible change that takes place in each test-tube.

What inference can be drawn from each observation?

What conclusion can be drawn from the effect produced by the addition of carbon tetrachloride? Account for these conclusions.

Write chemical equations to express the reactions that have taken place.

Is this a good test for a chloride?—a bromide?—an iodide? Explain.

3. To the second set of solutions, add small amounts of carbon tetrachloride, and observe. To each test-tube add 2 ml. of bromine water, shake, and observe.

What effect has carbon tetrachloride on a solution of (a) sodium chloride, (b) sodium bromide, and (c) sodium iodide?

What is the result of the addition of the bromine water?

What conclusion could be drawn from this experiment?

Could this test be used for an iodide? Write an equation to explain your answer.

4. To each test-tube of the third set of solutions, add 1 ml. of dilute nitric acid followed by 5 ml. of silver nitrate solution. Observe each reaction closely. Continue to watch the reactions, holding the test-tubes in sunlight if possible. Retain the test-tubes for Procedure 5.

Give all visible changes taking place in each test-tube.

Could the visible changes be used to identify separate samples of chlorides, bromides, and iodides? Explain.

Write chemical equations to express each of the reactions.

What is the effect of holding the test-tubes in sunlight?

Why is it necessary to add nitric acid? (TEXT p. 338)

5. To each test-tube used in Procedure 4, add 5 ml. of ammonium hydroxide solution, and shake well to mix its contents thoroughly.

What is the effect of the addition of ammonium hydroxide to each test-tube? Explain.

Describe how the addition of ammonium hydroxide may be considered as a confirmatory test.

QUESTIONS

1. What is the best test for a chloride?
2. What is the best test for a bromide?
3. What is the best test for an iodide?

EXPERIMENT 89. (*Demonstration*) The use of hydrogen fluoride to etch glass.

CAUTION: Hydrogen fluoride and its aqueous solution, hydrofluoric acid, are both very corrosive and must neither be inhaled nor allowed to come in contact with the skin.

1. Coat the inside of an evaporating dish with a generous layer of paraffin. Select a piece of window glass large enough to cover the dish adequately, and place an even coating of paraffin over its surface. With a pin or a fine stylus, scratch a design through the wax to the glass plate.

2. Place approximately 3 gm. of calcium fluoride in the paraffin-lined dish, and add sufficient concentrated sulphuric acid to make a thin paste.

3. Gently press the glass plate, wax-side down, on the evaporating dish, and set it in a fume cabinet or beside an exhaust vent for a time. After about an hour, scrape or melt the paraffin from the plate, and examine the surface of the glass.

Describe the appearance of the plate and account for the change that has taken place.

Why is it important to line the evaporating dish with paraffin?

How should hydrofluoric acid be stored in the laboratory?

UNIT 16

Compounds of Nitrogen

EXPERIMENT 90. To prepare and collect ammonia and study its properties.

PREPARATION AND COLLECTION

1. Place as much slaked lime as will lie on a ten-cent piece, in the palm of one hand. In the other hand, place an equal quantity of ammonium chloride. Holding the hands over the sink, rub the two chemicals together vigorously, and then **cautiously** note the odour of the gas produced.

What is the chemical name for slaked lime?

Describe the odour of the gas produced. Name the gas.

What three purposes are served by rubbing the hands vigorously?

Write the equation to represent the reaction.

2. Fill a Pyrex test-tube one-quarter full of ammonium chloride, add an equal quantity of calcium hydroxide, and shake so as to

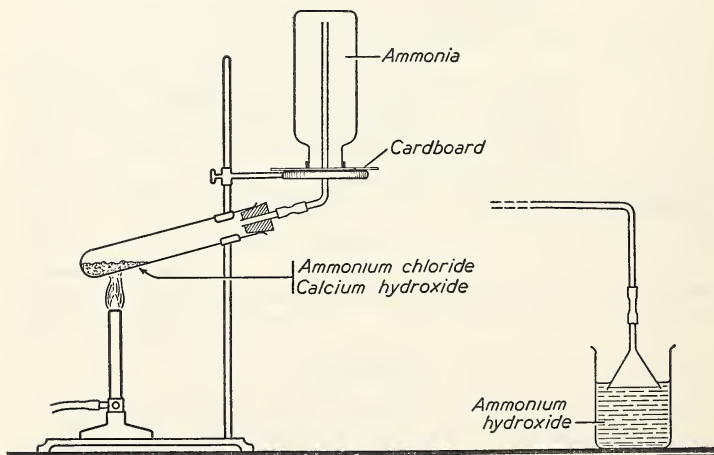


Fig. 75. The laboratory preparation and collection of ammonia.

mix the two solids thoroughly. Set up the apparatus as shown in Figure 75; heat the mixture gently, and collect three bottles of ammonia by the downward displacement of air. To determine when a bottle is filled, hold a piece of moist red litmus paper at the mouth. When the bottles are filled with ammonia, place them, mouth downward, on the table. Then, as illustrated in the diagram, attach the delivery tube containing the funnel to the generating test-tube, and just immerse the mouth of the funnel in water in a beaker.

What is the effect of ammonia on moist red litmus?

PROPERTIES OF AMMONIA

1. Hold a bottle of ammonia mouth down. Insert a blazing splint upward into the bottle, and observe.

What is the colour of ammonia gas?

What two properties of ammonia are illustrated by using the blazing splint?

Why is the bottle held mouth down?

Distinguish between "ammonia" and "ammonium".

2. Put four drops of concentrated hydrochloric acid into a gas bottle, cover with a glass plate, and shake. Place a bottle of ammonia, mouth upward, on the table. Invert the bottle containing the acid over the ammonia bottle, and remove the glass plate.

What vapour is given off by the concentrated acid?

Describe and account for the action that takes place within the bottles.

Which bottle shows the most pronounced change? Why?

Write the equation for the reaction.

3. Invert a bottle of ammonia in a beaker of water.

Describe and account for the change observed.

4. Test the solution of ammonium hydroxide, obtained in the preparation and collection of ammonia, with neutral litmus paper.

Write the equation that represents the reaction when ammonia "dissolves" in water.

What is the colour, odour, and effect on litmus of this solution?

5. Add 10 ml. of a solution of ammonium hydroxide to 10 ml. of a dilute solution of hydrochloric acid.

Write the equation for the reaction that takes place. Why does the ammonium chloride that is formed not appear as a white solid as it did in Procedure 2?

6. (**Demonstration**) Saturate a pile of a dozen pieces of filter paper with a concentrated solution of ammonium hydroxide, and place the pile on a glass plate at one end of the table. At the opposite end of

the table, saturate a pile of a dozen pieces of filter paper with concentrated hydrochloric acid. Now bring the two piles of filter paper together, one in each hand, and press them firmly against each other. Observe the reaction that takes place.

Describe and account for what happens.

What evidence is there that the reaction is exothermic?

EXPERIMENT 91. (*Demonstration*) An ammonia fountain.

1. Make a fine jet at one end of a piece of glass tubing about $1\frac{1}{2}$ feet long. Insert this tubing through one opening in a two-hole stopper, and a medicine dropper filled with water through the other opening. Allow the untapered end of the glass tubing to extend nearly to the bottom of a large beaker filled with a red litmus solution.

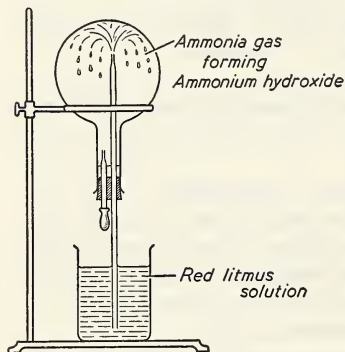


Fig. 76. An ammonia fountain.

2. Select a sturdy, round-bottomed flask and fill it with ammonia gas by the downward displacement of air. When the flask is filled, quickly insert the stopper into the inverted mouth, and support the flask in a retort ring as shown in Figure 76. Force a few drops of water out of the medicine dropper

into the flask, and observe the action that takes place.

Explain the formation and colour of the fountain.

QUESTIONS

1. Why is the heating of ammonium hydroxide not considered a laboratory method of preparing ammonia, even though it is a good source of ammonia?

2. Write nine equations to represent the production of ammonia gas, using three different ammonium salts and three different bases.

EXPERIMENT 92. (*Demonstration*) To prepare and collect nitric acid and study its properties.

PREPARATION AND COLLECTION

1. Arrange the apparatus as shown in Figure 77. Place two

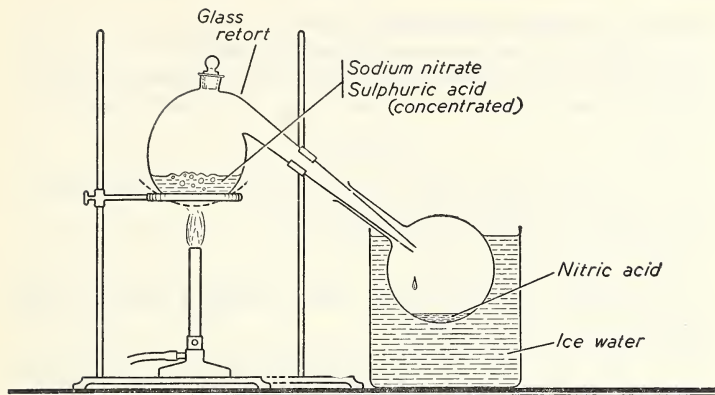


Fig. 77. The laboratory preparation and collection of nitric acid.

teaspoonfuls of sodium nitrate in the retort, and add sufficient concentrated sulphuric acid to cover the sodium nitrate.

2. Heat the mixture with a moderately hot flame, and observe the action in the retort and in the condensing flask.

3. When no more nitric acid collects in the condensing flask, remove the flame, and examine the contents of the retort and of the flask.

Describe the action that takes place in the retort.

What evidence is there that nitric acid has a low boiling point?

Describe the product remaining in the retort after the reaction.

Describe the nitric acid in the condensing flask.

Write the equation for the preparation of nitric acid.

PROPERTIES OF NITRIC ACID

CAUTION: Concentrated nitric acid is very corrosive and must be handled with great care.

Because the amount of nitric acid prepared in this experiment is small, it may be necessary to use additional concentrated nitric acid from a reagent bottle for some of the following experiments.

1. Place some white of egg in an evaporating dish, and add 2 ml. of concentrated nitric acid. Heat the evaporating dish, and observe the changes that take place.

Wash the excess nitric acid from the egg white, and then add 2 ml. of ammonium hydroxide, and observe any further changes.

Describe the effect of nitric acid on egg white.

What happens when ammonium hydroxide is added?

Egg white is a protein, and most proteins react in this way with concentrated nitric acid. The reaction is known as the xanthoproteic reaction.

2. Add one drop of concentrated nitric acid to one-third of a test-tube of water. Taste a drop of this solution. Note the effect of this solution on neutral litmus paper.

Describe the taste of the nitric acid solution.

What is the effect on litmus?

3. CAUTION: (Demonstration) When nitric acid boils, the liquid may spatter from the test-tube.

Heat 2 ml. of concentrated nitric acid in a Pyrex test-tube until it boils, and observe. Hold a piece of glowing charcoal, which has been formed on the end of a long, charred wooden splint, in the vapour above the boiling acid.

Describe all the changes that take place. Account for any colour change.

Write the equation for the decomposition of nitric acid by heat.

Account for the action of the glowing charcoal. Write the equation that represents the reaction.

What is the effect of the nitric acid vapour on the part of the wood that is not charred?

4. Drop a small piece of zinc into 2 ml. of dilute nitric acid in a test-tube, observe the reaction, and test the escaping gas with a blazing splint.

Describe the reaction in the test-tube. Is hydrogen produced?

5. Place a small piece of copper in each of two test-tubes. To one, add dilute nitric acid, and to the other, concentrated nitric acid, and observe.

What is observed in each test-tube?

Name the products of each reaction, and write the equations.

6. (Demonstration) Test a few drops of concentrated nitric acid with a solution of barium chloride to find out if it contains any trace of a sulphate. To 2 ml. of the same acid, add a piece of roll sulphur about the size of a bean, and boil the acid for several minutes. Test the resulting solution with barium chloride solution. Compare the resulting precipitate with the precipitate formed (if any) when the nitric acid was initially tested.

Describe and account for any reaction between barium chloride and the nitric acid tested. Why does commercial nitric acid often contain traces of a sulphate?

Describe and account for the observations noted when barium chloride solution is added after the nitric acid has been boiled with sulphur.

What important property of nitric acid is shown by this experiment?

Write the equation for the action of concentrated nitric acid on sulphur.

QUESTIONS

1. Why must nitric acid be handled with extreme care?
2. Why is sulphuric acid, rather than any other acid, used in the preparation of nitric acid?
3. Why is the apparatus used in the preparation of nitric acid made entirely of glass?
4. Explain why hydrogen is not produced when metals are added to either dilute or concentrated nitric acid. (TEXT p. 361)

EXPERIMENT 93. The properties of nitrates.

1. Examine samples of each of the following nitrates: sodium nitrate, potassium nitrate, lead nitrate, cupric nitrate, and barium nitrate.

Describe each of the nitrates selected.

2. Fill five Pyrex test-tubes one-quarter full of each nitrate, and heat each one intensely. Hold a glowing splint to the mouth of each test-tube in turn.

Describe and account for the action of the glowing splint.

Write a chemical equation for each reaction.

Describe and account for the action in each case.

3. Add a few crystals (or a solution) of sodium nitrate to 5 ml. of a freshly prepared, concentrated solution of ferrous sulphate in a test-tube, and shake to mix thoroughly. Clamp the test-tube at an angle of about 60° to the vertical, and slowly pour 2 ml. of concentrated sulphuric acid down the side of the tube so that the heavy acid will underlie the solution without much mixing.

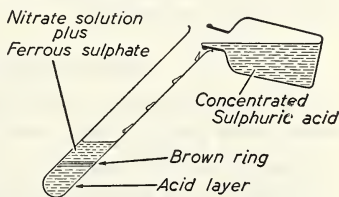


Fig. 78. Testing for nitrates (nitrate ion).

Describe the appearance at the junction of the acid and the solution.

Why is sulphuric acid added to the mixture in testing for a nitrate?

This procedure constitutes a delicate test for the nitrate ion.

EXPERIMENT 94. (*Demonstration*) To prepare and collect nitrous oxide and study its properties.

PREPARATION AND COLLECTION

CAUTION: Ammonium nitrate may explode if heated too rapidly. Nitrous oxide should not be prepared by students.

1. Place about 15 gm. of ammonium nitrate in a Pyrex test-tube, and arrange the apparatus as shown in Figure 79. Heat the ammonium nitrate gently until it melts, being very careful not to overheat.

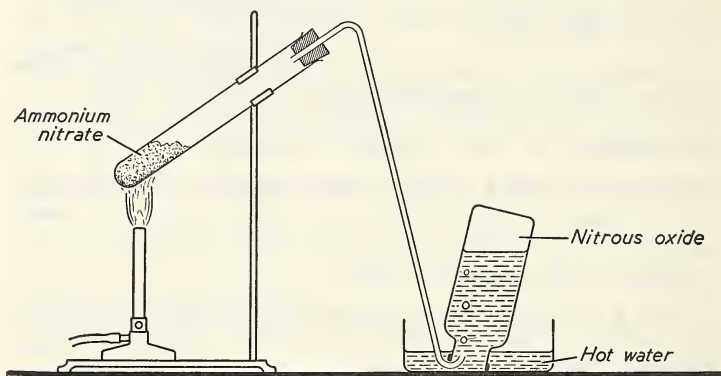


Fig. 79. The laboratory preparation and collection of nitrous oxide (laughing gas).

2. Collect three bottles of the gas by the downward displacement of hot water. Place a glass plate over the mouth of each bottle, and set them upright on the table.

PROPERTIES OF NITROUS OXIDE

1. Observe the colour and odour of nitrous oxide, and invert a bottle of the gas in a dish of cold water.

Describe the colour and odour of nitrous oxide.

Account for the action that takes place when a bottle full of the gas is inverted in cold water.

Why is hot water used in the original collection?

2. Insert a glowing splint into a bottle of the gas, and observe.

Does nitrous oxide support combustion?

Does nitrous oxide burn?

3. Place a piece of white phosphorus, the size of a grain of wheat,

on an asbestos-lined deflagrating spoon. Ignite the phosphorus, and lower it into a bottle of nitrous oxide.

Describe and account for the action that takes place.

QUESTIONS

1. Write the equation for the laboratory preparation of nitrous oxide.
2. Write the equation for the burning of (a) carbon, and (b) phosphorus, in nitrous oxide.
3. How can nitrous oxide be distinguished from oxygen?

EXPERIMENT 95. To prepare and collect nitric oxide and study its properties.

PREPARATION AND COLLECTION

1. Put approximately 20 gm. of copper turnings in a 250 millilitre flask, and set up the apparatus as in Figure 80. Through the thistle-tube, add sufficient dilute nitric acid to cover the copper. The flask may be warmed gently, but if allowed to stand, the reaction will proceed with increasing speed.

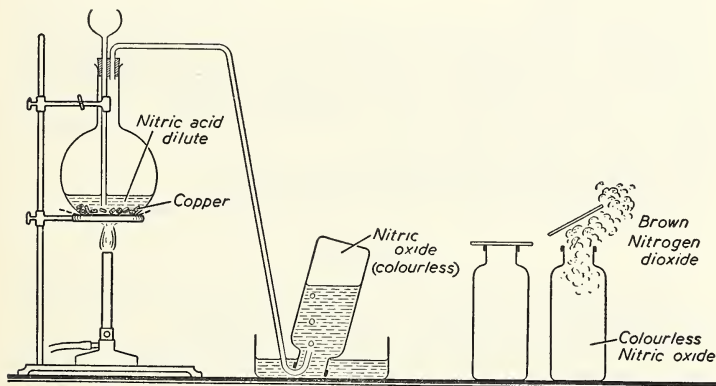


Fig. 80. The laboratory preparation and collection of nitric oxide.

2. Observe the changes that occur inside the flask. Collect four bottles of the gas by the downward displacement of water. Cover three of the bottles with glass plates, and place them upright on the table. Allow the fourth bottle to remain in the pneumatic trough until ready for use.

Describe the action of copper on dilute nitric acid.

What colour are the bubbles of gas that form in the liquid?

What is the colour of the solution?

What indication is there that the reaction is exothermic?

Describe and account for the appearance of the gas inside the flask.

What is the colour of the gas collected in the gas bottles? What has happened to the coloured gas inside the flask?

Write the equation for the preparation of nitric oxide.

PROPERTIES OF NITRIC OXIDE

1. Observe the colour of nitric oxide, and remove the glass plate from one of the bottles.

What is the colour of nitric oxide?

What happens when the glass plate is removed?

Why is it impossible to note the odour of nitric oxide?

Name the gas produced when nitric oxide is exposed to air.

Write the equation for the reaction that takes place when nitric oxide is exposed to air.

2. (**Demonstration**) Lower (a) a blazing splint, and (b) burning white phosphorus in an asbestos-lined deflagrating spoon, into separate bottles of nitric oxide.

Does the gas burn?

Does nitric oxide support the combustion of (a) wood, (b) phosphorus?

3. Allow a few bubbles of oxygen from an oxygen generator to enter the bottle remaining in the pneumatic trough. Observe the changes that take place, and continue to add oxygen, a few bubbles at a time, until no further changes occur.

Describe and account for the changes that take place.

Write the equation for the reaction when the two gases mix.

QUESTIONS

1. How does the decomposition temperature of nitric oxide compare with that of nitrous oxide? Explain your answer.

2. How could nitric oxide be used as a test for oxygen? Explain with the aid of an equation.

EXPERIMENT 96. The preparation of nitrogen dioxide and a study of its properties.

To a few copper turnings in a test-tube, add 2 ml. of concentrated nitric acid, and observe the action that takes place. Lower a blazing splint into the test-tube.

*What is the colour of the bubbles of gas that appear in the nitric acid?
What is the colour of the solution?*

What change takes place in the colour of the solution as the reaction continues? Explain.

Does nitrogen dioxide burn? Does it support combustion?

Is nitrogen dioxide more dense or less dense than air?

What method could be used to collect a bottle of the gas? Why?

Write the equation for the laboratory preparation of nitrogen dioxide.

UNIT 17

Phosphorus and its Compounds

EXPERIMENT 97. (*Demonstration*): The allotropic modifications of phosphorus.

1. Using tongs, remove a stick of white phosphorus from a stock bottle, and immediately place it in a shallow dish containing sufficient water to cover it completely. Using a sharp knife and keeping the phosphorus under water, cut off several small pieces approximately the size of a match head.

Why is it essential to keep the phosphorus under water while it is being cut?

Describe the appearance of the phosphorus on a freshly cut surface.

2. Using forceps, place one of the small pieces of phosphorus on a dry filter paper, and then transfer it immediately to a dry porcelain crucible. Observe the effect of exposing the phosphorus to air for a short time, and then touch it with the end of a glass rod that has been warmed in a Bunsen flame.

Why is it a good precaution to dry the phosphorus?

Does any change occur when the phosphorus is exposed to air? Explain.

Describe the changes that take place when white phosphorus burns.

What is the colour of the deposit remaining in the crucible? What is this deposit?

Write chemical equations for the burning of white phosphorus in air.

3. Using forceps, place one of the small pieces of phosphorus on a glass plate. Darken the room as completely as possible, and observe the appearance of the phosphorus.

Describe and account for the appearance of the phosphorus.

4. **CAUTION:** Turn off all burner flames before using carbon disulphide. It is a highly flammable liquid.

Pour 2 ml. of carbon disulphide into a test-tube, and add two of the small pieces of white phosphorus. Leaving the mouth uncovered, gently agitate the test-tube, and observe any effect on the pieces of phosphorus.

Is white phosphorus soluble in carbon disulphide?

5. Support a piece of filter paper on a wire gauze in the sink, and saturate it with a sufficient quantity of the phosphorus—carbon disulphide solution. Stand back, and observe the action that takes place.

Describe and account for the action that occurs.

6. Into a crucible, place as much red phosphorus as will lie on the point of a knife blade, and observe for a short time. Then darken the room, and compare the appearance of the red phosphorus with that of white phosphorus under similar conditions. Touch the red phosphorus with the end of a glass rod that has been warmed in a Bunsen flame. Then heat the end of the glass rod intensely, and again touch it to the red phosphorus.

Is there any noticeable change in the red phosphorus when it is exposed to air?

Does red phosphorus phosphoresce?

Does red phosphorus ignite when touched with (a) a warm glass rod, (b) a hot glass rod?

Describe the burning of red phosphorus.

Write chemical equations for the burning of red phosphorus in air.

7. Place a small amount of red phosphorus in carbon disulphide in a test-tube. Leaving the mouth uncovered, gently agitate the test-tube and observe the effect.

Is red phosphorus soluble in carbon disulphide?

EXPERIMENT 98. (*Demonstration*): The oxides and acids of phosphorus.

1. Line a deflagrating spoon with asbestos paper, and in it, place a small piece of white phosphorus. Warm the spoon gently in a burner flame, and then lower it into a bottle filled with oxygen. Save this bottle for Procedure 3.

Describe and account for the action that takes place.

2. Repeat Procedure 1, using red phosphorus instead of white phosphorus. Heat the deflagrating spoon until the red phosphorus ignites, and then quickly lower the spoon into a bottle filled with oxygen. Save this bottle for Procedure 3.

Describe and account for the action that takes place.

3. After removing the deflagrating spoons from the bottles used in Procedures 1 and 2, pour 10 ml. of water into each bottle. Place a glass plate over the mouth of each bottle and shake. To each bottle add a piece of red litmus paper and a piece of blue litmus paper.

Name and account for the substances contained in each bottle after the phosphorus stops burning.

What evidence is there that these substances are soluble in water?

Describe and account for the action on litmus.

Write equations to represent, (a) the burning of white phosphorus in oxygen, (b) the burning of red phosphorus in oxygen, (c) the dissolving of the oxides in water.

Name the solution formed when phosphoric oxide is dissolved in water.

Name the solution formed when phosphorous oxide is dissolved in water.

4. Repeat Procedures 1, 2, and 3, using bottles filled with air instead of oxygen.

Account for the differences observed when air is used in place of oxygen.

Name and account for the substances produced when (a) white phosphorus burns in air, (b) red phosphorus burns in air, (c) the oxides dissolve in water.

Write equations to represent the burning of phosphorus in air, and name the acids formed when these oxides dissolve in water.

UNIT 18

Calcium and Magnesium and their Compounds

EXPERIMENT 99. To prepare a sample of calcium carbonate (precipitated chalk).

1. Mix together 50 ml. of a saturated solution of calcium chloride, and 50 ml. of a saturated solution of sodium carbonate in a 250 millilitre beaker, stir, and observe.

2. Filter the mixture. Rinse the beaker with water to make sure that all of the contents are transferred to the filter paper.

3. Wash the residue on the filter paper several times with water, and allow it to dry thoroughly. Then scrape it onto a watch-glass and note its appearance.

4. To a small portion of this residue in a test-tube, add 2 ml. of dilute hydrochloric acid. Hold a drop of limewater on the end of a glass tube in the mouth of the test-tube, and observe.

Describe what happens when the two solutions are mixed.

What is the colour of the precipitate?

Why was the residue washed with water?

Describe the appearance of the dried residue.

Describe and account for the action of the acid on the residue.

Write the chemical equation to represent this method of making calcium carbonate.

What property of calcium carbonate makes possible its manufacture by this method?

EXPERIMENT 100. To prepare quicklime from calcium carbonate.

1. Moisten a piece of marble (calcium carbonate) about the size of a bean, and touch it to a piece of red litmus paper.

2. Put the piece of marble into a clean crucible, and heat it intensely for twenty minutes.

3. Allow the crucible and contents to cool to room temperature, and examine.

4. Touch one edge of the material in the crucible with a piece of moist red litmus paper.

5. Allow a drop of water to fall on the material in the crucible, and observe. Add a few more drops of water, and note if there is any further change. Then add 5 ml. of water, stir, and filter the mixture. Catch the filtrate in a test-tube. Blow your breath through the filtrate and observe.

Does the marble chip have any effect on moist litmus?

Describe the changes observed in the marble when it is strongly heated.

Name and describe the product remaining in the crucible after it cools.

Does the product remaining in the crucible have any effect on moist red litmus?

Write the equation for the preparation of quicklime.

Describe the action that takes place when a drop of water is added to the substance in the crucible.

What is observed when the breath is blown through the filtrate? Name the filtrate.

EXPERIMENT 101. The slaking of quicklime.

CAUTION: When quicklime is slaked, a minor explosion may occur.

Place a piece of quicklime, the size of a walnut, in an evaporating dish. Protecting your eyes with a glass plate, drop 4 drops of water on the lime and observe. Continue to add water, a few drops at a time, as long as any change takes place. Finally, add about 20 ml. of water, stir the mixture, and filter some of it into a test-tube. Blow your breath through the filtrate.

Describe all the changes that take place in the quicklime when water is slowly added.

What evidence is there that the slaking of quicklime is an exothermic reaction?

Describe the slaked lime prepared, and compare it with the quicklime used in its preparation.

Write the equation for the slaking of quicklime.

What is the name of a water solution of slaked lime?

What two changes must take place before quicklime can be changed to limewater?

EXPERIMENT 102. The preparation of bleaching powder.

1. Fill a test-tube one-third full of calcium hydroxide. Support the test-tube in a horizontal position, and spread the calcium hydroxide evenly along the tube.

2. Place a piece of glass tubing, leading from a chlorine generator, into the test-tube so that the end of the tubing is near the closed end of the test-tube. Slowly pass chlorine into the test-tube for at least ten minutes.

3. Empty the freshly made bleaching powder from the test-tube into a beaker, and note its odour. Add water so as to make a thin paste, and divide the paste into two parts.

4. To one part of the mixture, add 5 ml. of dilute sulphuric acid. Cautiously note the odour of the gas produced, and hold a piece of moist blue litmus paper in the gas.

5. Place a piece of fruit-basket netting, or some similar coloured cloth, in the second portion of the paste, and after it is well coated, drop it into a beaker containing dilute sulphuric acid. Remove the cloth, wash it, and note any colour change.

Is there any noticeable change in the slaked lime when chlorine is passed over it?

Describe and account for the effect of sulphuric acid on bleaching powder.

Explain the action on the moist litmus paper.

Describe and account for the colour changes in the cloth.

Write the equation for the preparation of bleaching powder from slaked lime and chlorine.

Write the equation for the liberation of chlorine from bleaching powder.

EXPERIMENT 103. The setting of plaster of Paris.

1. Examine a lump of gypsum and a sample of plaster of Paris.

From a comparison of physical properties, how would you distinguish gypsum from plaster of Paris?

2. Place two teaspoonfuls of plaster of Paris in an evaporating dish, and add water, a few drops at a time, stirring the mixture constantly. When the mixture is in the form of a thick paste, transfer it to a glass plate which has been lightly coated with vaseline (petroleum jelly).

3. Into the plaster of Paris, press a coin (or other object) which has been lightly coated with vaseline. Scrape away any projections of the plaster which, after hardening, might interfere with the removal of the coin. After the plaster of Paris has hardened or "set", carefully pry the coin from its position, and observe its imprint.

What useful purpose is served by the vaseline?

Why is the imprint of the coin perfect in every detail?

What causes the plaster of Paris to harden? Write the equation that represents the setting of plaster of Paris.

EXPERIMENT 104. Water of temporary hardness.

1. Fill a test-tube one-quarter full of distilled water, and add one drop of soap solution. Cover the mouth of the tube with your thumb, and shake it vigorously back and forth ten times. Observe the effect of the soap on the distilled water. Continue to add soap solution, one drop at a time, shaking after each addition. Count the number of drops of soap solution required to produce a permanent lather. A permanent lather is one that will last for at least one minute.

What effect is produced when soap solution is added to distilled water?

How many drops of soap solution are required to produce a lather that lasts for at least one minute?

Why is distilled water considered to be "soft"? (TEXT p. 380)

2. Half fill a large test-tube with limewater, and slowly bubble carbon dioxide through the limewater until the white precipitate that forms, dissolves. The water now contains a solution of calcium hydrogen carbonate (calcium bicarbonate), and is said to have "temporary hardness". Divide the clear solution of calcium bicarbonate into four equal parts by pouring it into four separate test-tubes.

Write equations to represent (a) the formation of the white precipitate, and (b) the "dissolving" of the precipitate.

3. To one of the test-tubes from Procedure 2, add soap solution, one drop at a time. Count the drops as they are added, and after adding each drop, hold the thumb over the mouth of the test-tube, and shake vigorously ten times. Continue adding the soap solution until a lather that lasts for at least one minute is produced. Record the number of drops of soap required to produce this permanent lather.

Describe the action of the soap on the hard water.

When does the soap produce a lather with the water? How many drops of soap solution were required to produce this lather?

Explain how the addition of soap softens water of temporary hardness.

4. Hold a second test-tube containing water of temporary hardness in the flame of a Bunsen burner, and heat slowly until it comes to the boiling point. Observe the changes that take place, and then cool the solution to room temperature. Add soap solution to the test-tube, as in Procedure 3, and count the number of drops required to produce a permanent lather.

Describe and account for the changes that take place when water of temporary hardness is boiled. Write the equation to represent the reaction.

How many drops of soap solution were required to produce a permanent lather?

5. To one of the remaining test-tubes, add a crystal of washing soda, and shake. Observe any changes in the contents of the test-tube, and test with a soap solution, as in Procedure 3.

Describe and account for any changes that take place.

Write the equation for the action of washing soda on water of temporary hardness.

How many drops of soap solution were required to produce a permanent lather?

How does washing soda act in softening water of temporary hardness?

6. To the remaining test-tube, add 5 ml. of limewater and observe any changes. Test with soap as in Procedure 3.

Describe and account for any changes observed.

Write the equation for the reaction.

How many drops of soap solution were required to produce a permanent lather? How does this number compare with the number of drops required in each of the above Procedures?

Does the addition of limewater soften water of temporary hardness as efficiently as (a) boiling, and (b) the addition of washing soda?

QUESTIONS

1. Write equations to show how water of temporary hardness may be prepared in the laboratory.

2. What is meant by "water of temporary hardness"?

3. From the above experiments, name four ways by which temporary hardness may be removed from water. Which method is the most effective? Why?

EXPERIMENT 105. Water of permanent hardness.

1. Fill a test-tube one-quarter full of distilled water, and add one drop of soap solution. Cover the mouth of the tube with your thumb, and shake it vigorously back and forth ten times. Observe the effect of the soap on the distilled water. Continue to add soap solution, one drop at a time, shaking after each addition and counting the number of drops, until a permanent lather is produced.

What effect is produced when soap solution is added to distilled water?

How many drops of soap solution were required to produce a permanent lather?

2. Fill each of two test-tubes one-quarter full of distilled water. To one test-tube, add a piece of calcium chloride the size of a pin head,

and to the other, add a similar sized piece of magnesium sulphate. Add one drop of soap solution to each test-tube, and shake vigorously ten times. Continue to add soap solution, one drop at a time, shaking each time and counting the number of drops, until a permanent lather is produced. Observe all the changes that take place.

Describe and account for the change that takes place when soap is added to a distilled water solution of (a) calcium chloride, and (b) magnesium sulphate.

How many drops of soap solution were required to produce a permanent lather in each case?

In what respect is permanent hardness of water similar to temporary hardness?

What general reaction takes place whenever water is softened?

Does the addition of soap soften permanently hard water? Explain.

3. Fill a test-tube one-quarter full of distilled water, and add a piece of calcium chloride the size of a pin head. Heat the solution to the boiling point, allow the test-tube to cool, and test the solution with soap solution, as in Procedure 2.

How many drops of soap solution were required to produce a permanent lather?

Does boiling the calcium chloride solution produce any visible change?

What effect has boiling water of permanent hardness on the hardness of the water?

4. Fill a test-tube one-quarter full of distilled water, and add a piece of magnesium chloride the size of a pin head. To this solution, add a crystal of washing soda the size of a bean, shake the test-tube, and observe any changes that take place. Test the resulting mixture with soap solution to determine the number of drops required to produce a permanent lather.

What visible change takes place when a washing soda crystal is added to the magnesium chloride solution?

Account for the action of the sodium carbonate, and write the equation to express the reaction that takes place.

How many drops of soap solution were required to produce a permanent lather?

Why is washing soda an effective water softener?

QUESTIONS

1. List all the substances whose solutions cause water to be hard.
2. What is a common method of detecting hardness in water?
3. Discuss the relative merits of the terms: *temporary* or *carbonate* hardness; *permanent* or *non-carbonate* hardness.

4. What advantage is gained by using chemical water softeners when softening water?

UNIT 19

Rates of Chemical Reactions

EXPERIMENT 106. The factors affecting the rates of chemical reactions.

NATURE OF THE REACTANTS

1. Fill each of two large Pyrex test-tubes ($1" \times 7\frac{1}{2}"$) one-quarter full of dilute hydrochloric acid, and stand them in a rack. Then, weigh out (a) 200 mg. of powdered magnesium, and (b) 200 mg. of powdered iron, on separate pieces of paper.

2. Have a watch, equipped with a second hand, in readiness to time the reactions. Then add the magnesium powder to one of the tubes containing hydrochloric acid, and determine accurately the time required for the reaction to go to completion. As the reaction proceeds, insert a blazing splint into the upper part of the tube.

Time required for the completion of this reaction . . . ____ seconds.

What gas is produced in this reaction?

Write the chemical equation for this reaction.

3. Add the powdered iron to the remaining test-tube containing hydrochloric acid, and determine the time required for the reaction to go to completion. As the reaction proceeds, insert a blazing splint into the upper part of the tube.

Time required for the completion of this reaction . . . ____ seconds.

What gas is produced in this reaction?

Write the chemical equation for this reaction.

Account for the different times required for these reactions to go to completion.

TEMPERATURE

1. Select two small pieces of mossy zinc, each weighing approximately 1 gm., (or pick two small pieces of about the same volume). Fill each of two large test-tubes one-quarter full of dilute hydrochloric acid, and add a piece of zinc to each tube. Observe the vigour of the reaction in both tubes.

2. While the reaction continues in both tubes, heat one of them gently. Compare the rates of reaction in the two test-tubes.

What is the only condition that differs in the two test-tubes?

What is the effect of this factor on the rate of the reaction?

Write the equation for this reaction.

CONCENTRATION

1. Using two large test-tubes (1" $\times$ 7½"), add water to a height of one inch in one tube and five inches in the other. Then add 5 ml. of concentrated hydrochloric acid to each tube.

2. Select two small pieces of mossy zinc, each weighing approximately 1 gm., (or select two small pieces of about the same volume). Simultaneously, add one of the pieces of zinc to each of the acid solutions prepared in Procedure 1. Compare the vigour of the two reactions.

Account for the different times required for these reactions to go to completion.

STATE OF SUBDIVISION

1. Place a 200 milligram calcium carbonate tablet in a large test-tube. Crush a second tablet to a fine powder on a piece of folded glazed paper, and tap it gently into a second large test-tube. (Calcium carbonate tablets of this mass may be secured from Starkman Chemists, 459 Bloor Street West, Toronto 5, Ontario, or possibly from a local druggist.)

2. In turn, half fill each test-tube with dilute hydrochloric acid of the same concentration. Observe the vigour and duration of the action in each tube.

Account for the difference in the rates of the reactions.

3. Into two separate test-tubes, place lead nitrate crystals to a depth of one-quarter inch. Into two other test-tubes, place potassium iodide crystals to a depth of one-quarter inch.

Describe the appearance of these two chemicals.

4. To one of the test-tubes containing lead nitrate crystals, add the potassium iodide crystals from one of the other test-tubes. Hold the thumb over the end of the tube, and shake the two solids together vigorously. Look for any evidence of a chemical reaction taking place.

Account for any colour change within the test-tube.

5. To the two remaining test-tubes containing crystalline lead nitrate and crystalline potassium iodide respectively, add water to fill each one-quarter full.

Describe the appearance of each solution.

6. Pour the potassium iodide solution into the solution of lead nitrate, and observe the reaction.

Account for any change in appearance within the test-tube.

Write the chemical equation for this reaction.

7. Account for the difference in the rates of the reactions in Procedure 4 and Procedure 6.

CATALYSIS

1. (*Demonstration*) Place one-half teaspoonful of zinc dust in a mortar, and add a similar volume of powdered iodine crystals. Grind the two substances thoroughly, and transfer the mixture to the lid of a tin can. Using an atomizer or a wash bottle, direct a fine spray of water against the mounded material.

Was there any evidence of reaction when the two solids were ground together?

What was observed when water was added?

What role did the water play in this experiment?

Write the equation for this reaction.

2. Half fill each of two large test-tubes with a 3% solution of hydrogen peroxide, and stand them in a rack. Insert a glowing splint into the mouth of each one.

Is there any effect on a glowing splint?

3. Using a flattened wood splint, slowly add small amounts of manganese dioxide to one of the tubes containing the hydrogen peroxide solution until a noticeable change is produced. Again insert a glowing splint into the mouth of each tube.

What do you observe?

What role did the manganese dioxide play in this experiment?

Write the equation for this reaction.

4. Place equal masses (approximately 1 gm.) of *pure* granulated zinc in each of two test-tubes.

5. Add 5 ml. of dilute sulphuric acid to one test-tube, and place it in a rack. Examine it from time to time.

Is there any evidence of reaction?

6. Add 5 ml. of cupric sulphate solution to the other test-tube containing pure zinc. Note any change in the appearance of the zinc and the colour of the solution. After two or three minutes, carefully pour off the solution remaining, and discard it into the sink. Add 5 ml. of dilute sulphuric acid, and test any gas produced by inserting a blazing splint in the upper part of the tube.

What colour changes are observed in the test-tube after the cupric sulphate solution is added?

Write an equation for the action of cupric sulphate solution on pure zinc.

Account for the colour changes seen in the test-tube.

Account for the difference in the rates of reaction in Procedures 5 and 6.

RADIANT ENERGY

1. Half fill two separate test-tubes with a solution of sodium chloride, and fill two other test-tubes one-quarter full of a solution of silver nitrate.

Describe the appearance of each solution.

2. In a darkened room, add each silver nitrate solution to each sodium chloride solution.

What is the appearance of the mixture immediately after the silver nitrate has been added?

Write a chemical equation to represent the reaction between solutions of silver nitrate and sodium chloride.

3. Place one of the mixtures, produced in Procedure 2, in a dark cupboard, and place the other in bright sunlight. After an interval of several minutes, hold the two test-tubes side by side, and compare their appearance.

What is the only condition that differs in the two test-tubes?

How does this experiment illustrate that radiant energy affects the rate of this reaction?

Appendix

A. IMPORTANT UNITS IN THE METRIC SYSTEM

Length

10 millimetres (mm.)	=	1 centimetre (cm.)
10 centimetres	=	1 decimetre (dm.)
10 decimetres	=	1 metre (m.)
1000 metres	=	1 kilometre (km.)

Mass (Weight)

10 milligrams (mg.)	=	1 centigram (cg.)
10 centigrams	=	1 decigram (dg.)
10 decigrams	=	1 gram (gm.)
1000 grams	=	1 kilogram (kg.)

Volume

1000 millilitres (ml.)	=	1 litre (l.)
1 millilitre	=	0.001 litre
1 millilitre	=	1.000028 cubic centimetres (cc.)

While a millilitre (ml.) and a cubic centimetre (cc.) have not exactly the same value, the difference is so slight that for practical purposes these two units may be considered to be identical.

B. RELATIONSHIP OF SOME ENGLISH AND METRIC UNITS

1 inch (in.)	=	2.54 centimetres
1 metre	=	39.37 inches
1 kilometre	=	0.62 miles
1 pound (lb.) Avoirdupois	=	453.6 grams
1 ounce (oz.) Avoirdupois	=	28.35 grams
1 kilogram	=	2.20 pounds
1 litre	=	0.88 quarts (British)
1 cubic foot (cu. ft.)	=	28.32 litres

C. GENERAL RULES FOR SOLUBILITY (In Water)

1. All *sodium*, *potassium*, and *ammonium* compounds are soluble.
2. All *nitrates*, *chlorates*, and *acetates* are soluble.
3. All *chlorides*, *bromides*, and *iodides* are soluble except those of *silver*, *lead*, and *mercury* (mercurous).
4. All *sulphates* are soluble except those of *barium*, *lead*, *calcium* (slightly), *silver* (slightly).
5. All *hydroxides* are insoluble except those of *sodium*, *potassium*, *ammonium*, *calcium*, and *barium*.
6. All metallic *sulphides* are insoluble except those of *sodium*, *potassium*, *ammonium*, *calcium*, and *barium*.
7. All *carbonates*, *phosphates*, and *silicates* are insoluble except those of *sodium*, *potassium*, and *ammonium*.
8. All metallic *oxides* are insoluble, but a few such as the oxides of *potassium*, *sodium*, *calcium*, and *barium*, **react** with water forming soluble hydroxides.

There are a few exceptions to these general rules, but the exceptions are not important to a beginner in the study of chemistry.

D. ACTIVITY TABLE (Some of the common metals)

1. Potassium	9. Tin
2. Sodium	10. Lead
3. Barium	11. Hydrogen
4. Calcium	12. Copper
5. Magnesium	13. Mercury
6. Aluminum	14. Silver
7. Zinc	15. Platinum
8. Iron	16. Gold

The relative activity with which metals displace *hydrogen* from compounds such as water and dilute acids is shown by this table. Metals *below* hydrogen will not displace hydrogen from these compounds.

E. TABLE OF SOLUBILITIES

(In cold water)

S = soluble in water

M = slightly or moderately soluble in water

I = insoluble or nearly insoluble in water

— = nonexistent compound, or decomposed by water

R = reacts chemically with water forming a base

	Acetate	Bromide	Carbonate	Chlorate	Chloride	Hydroxide	Iodide	Nitrate	Oxide	Phosphate	Sulphate	Sulphide	Sulphite
Aluminum	S	S	—	S	S	I	S	S	I	I	S	—	—
Ammonium	S	S	S	S	S	S	S	S	—	S	S	S	S
Barium	S	S	I	S	S	M	S	S	R	I	I	S	M
Calcium	S	S	I	S	S	M	S	S	R	I	M	M	M
Copper (—ic)	S	S	I	S	S	I	S	S	I	I	S	I	—
Iron (—ous)	S	S	I	—	S	I	S	S	I	I	S	I	M
Iron (—ic)	S	S	—	—	S	I	—	S	I	I	S	I	—
Lead	S	M	I	S	I	I	I	S	I	I	I	I	I
Magnesium	S	S	I	S	S	M	S	S	R	I	S	I	M
Manganese	S	S	I	—	S	I	—	S	I	I	S	I	—
Mercury (—ous)	S	I	I	S	I	—	I	S	I	I	M	—	—
Mercury (—ic)	S	M	I	S	S	—	I	S	I	I	S	I	—
Potassium	S	S	S	S	S	S	S	S	R	S	S	S	S
Silver	M	I	I	S	I	—	I	S	I	I	M	I	S
Sodium	S	S	S	S	S	S	S	S	R	S	S	S	S
Strontium	S	S	I	S	S	M	S	S	R	I	I	M	—
Zinc	S	S	I	S	S	I	S	S	I	I	S	I	I

F. GENERAL PROPERTIES OF METALS AND NON-METALS**Metals****PHYSICAL PROPERTIES**

1. Good conductors of heat.
2. Good conductors of electricity.
3. Greyish metallic lustre, except copper (red) and gold (yellow).
4. Malleable. (Gold may be pounded into sheets 0.0001 cm. thick.)
5. Ductile. (Wire may be made weighing 0.000005 gm./cm.)

CHEMICAL PROPERTIES

1. All form oxides either by direct combination or by some intermediate reaction.
2. Metallic oxides (if they react with water) are basic anhydrides.
3. The more active metals (those above **Hydrogen** in the activity series) react with dilute acids, liberating hydrogen.
4. The more active metals react with water, forming hydrogen and a base.
5. Metals react with non-metals.
6. Metals form positive ions.

Non-Metals**PHYSICAL PROPERTIES**

1. Poor conductors of heat.
2. Poor conductors of electricity, except graphite and selenium.
3. Vary considerably in colour; sulphur (yellow), phosphorus (white or yellow or red), chlorine (greenish-yellow), bromine (brownish-red), iodine (steel-grey with violet vapour).
4. Not malleable.
5. Not ductile.

CHEMICAL PROPERTIES

1. All form oxides either by direct combination or by some intermediate reaction.
2. Non-metallic oxides (if they react with water) are acid anhydrides.
3. Only the most active non-metals react with water, and when they do, they form acids, and liberate oxygen.
4. Non-metals react with metals.
5. Non-metals form negative ions.

G. PRESSURE OF WATER VAPOUR, OR AQUEOUS TENSION*(In millimetres of mercury)*

TEMPER- ATURE	PRESSURE	TEMPER- ATURE	PRESSURE	TEMPER- ATURE	PRESSURE
0°C.	4.6 mm.	11°C.	9.8 mm.	22°C.	19.8 mm.
1°	4.9 mm.	12°	10.5 mm.	23°	21.0 mm.
2°	5.3 mm.	13°	11.2 mm.	24°	22.3 mm.
3°	5.7 mm.	14°	12.0 mm.	25°	23.7 mm.
4°	6.1 mm.	15°	12.8 mm.	26°	25.1 mm.
5°	6.5 mm.	16°	13.6 mm.	27°	26.7 mm.
6°	7.0 mm.	17°	14.5 mm.	28°	28.3 mm.
7°	7.5 mm.	18°	15.5 mm.	29°	29.9 mm.
8°	8.0 mm.	19°	16.5 mm.	30°	31.7 mm.
9°	8.6 mm.	20°	17.5 mm.	50°	92.3 mm.
10°	9.2 mm.	21°	18.6 mm.	100°	760.0 mm.

When a gas is collected over water, the true pressure exerted by the gas itself is obtained by subtracting from the observed barometric pressure the pressure due to water vapour at the temperature of measurement.

H. DENSITY OF GASES*(At 0°, and 760 mm. pressure)*Based on *approximate* atomic weights

Name	Form- ula	Mol. Wt.	Density (gm./litre)
Acetylene	C_2H_2	26	1.16
Air			1.29
Ammonia	NH_3	17	0.76
Argon	A	39.94	1.78
Carbon dioxide	CO_2	44	1.96
Carbon monoxide	CO	28	1.25
Chlorine	Cl_2	71	3.17
Fluorine	F_2	38	1.70
Helium	He	4	0.18
Hydrogen	H_2	2	0.09
Hydrogen bromide	HBr	81	3.61
Hydrogen chloride	HCl	36.5	1.63

Name	Form- ula	Mol. Wt.	Density (gm./litre)
Hydrogen iodide	HI	128	5.71
Hydrogen sulphide	H₂S	34	1.52
Krypton	Kr	83.7	3.73
Methane	CH₄	16	0.71
Neon	Ne	20	0.90
Nitric oxide	NO	30	1.34
Nitrogen	N₂	28	1.25
Nitrogen dioxide	NO₂	46	2.05
Nitrous oxide	N₂O	44	1.96
Oxygen	O₂	32	1.43
Sulphur dioxide	SO₂	64	2.86
Xenon	Xe	131	5.85

I. APPROXIMATE ATOMIC WEIGHTS OF THE COMMON ELEMENTS

NAME	SYMBOL	ATOMIC WEIGHT	NAME	SYMBOL	ATOMIC WEIGHT
Aluminum	Al	27	Lead	Pb	207
Antimony	Sb	122	Magnesium	Mg	24
Arsenic	As	75	Manganese	Mn	55
Barium	Ba	137	Mercury	Hg	200
Bromine	Br	80	Nitrogen	N	14
Calcium	Ca	40	Oxygen	O	16
Carbon	C	12	Phosphorus	P	31
Chlorine	Cl	35.5	Potassium	K	39
Copper	Cu	63.5	Silver	Ag	108
Fluorine	F	19	Sodium	Na	23
Gold	Au	197	Sulphur	S	32
Hydrogen	H	1	Tin	Sn	119
Iodine	I	127	Zinc	Zn	65
Iron	Fe	56			

J. PREPARATION OF LABORATORY REAGENTS

Acetic Acid (dilute)	: 1 vol. glacial with $2\frac{1}{2}$ vols. water.
Hydrochloric acid (conc.)	: use the C.P. acid of commerce.
Hydrochloric acid (dilute)	: equal vols. of C.P. acid and water.
Nitric acid (conc.)	: use the C.P. acid of commerce.
Nitric acid (dilute)	: 1 vol. C.P. acid with 2 vols. water.
Sulphuric acid (conc.)	: use the C.P. acid of commerce.
Sulphuric acid (dilute)	: 1 vol. C.P. acid with 5 vols. water.
Ammonium hydroxide (conc.)	: use the C.P. product of commerce.
Ammonium hydroxide (dilute)	: 1 vol. C.P. product with $1\frac{1}{2}$ vols. water.
Potassium hydroxide (conc.)	: use a saturated solution.
Potassium hydroxide (dilute)	: dissolve 336 gm. of solid potassium hydroxide in water, and dilute to 1 litre.
Sodium hydroxide (conc.)	: use a saturated solution.
Sodium hydroxide (dilute)	: dissolve 240 gm. of solid sodium hydroxide in water, and dilute to 1 litre.
Silver nitrate solution	: dissolve 34 gm. of solid silver nitrate in water and dilute to 1 litre.

K. SULPHURIC ACID—DICHROMATE CLEANING SOLUTION

Cautiously (*Technique 3*) pour 200 ml. of concentrated sulphuric acid into 150 ml. of cold water, stirring continuously, and without further heating, saturate the hot solution with powdered potassium dichromate.

To clean glass vessels, fill them with this cold solution, and allow to stand overnight or longer.

L. SOME COMMON NAMES OF SUBSTANCES AND THEIR FORMULAE

COMMON NAME	FORMULA
Alum	$K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24 H_2O$
Ammonia water	NH_4OH
Aqua regia	HCl (4 vol.), HNO_3 (1 vol.)
Baking powder	$NaHCO_3$ + starch + an acid-acting ingredient
Baking soda	$NaHCO_3$
Bleaching powder	$CaOCl_2$

COMMON NAME	FORMULA
Blue vitriol (bluestone)	$\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$
Bone black	C
Borax	$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10 \text{H}_2\text{O}$
Caustic potash	KOH
Caustic soda	NaOH
Chile saltpetre	NaNO_3
Chloride of lime	CaOCl_2
Coke	C
Common salt	NaCl
Cream of tartar	$\text{KHC}_4\text{H}_4\text{O}_6$
Dextrose (glucose)	$\text{C}_6\text{H}_{12}\text{O}_6$
Diamond	C
Dry ice	CO_2 (solid)
Epsom salt	$\text{MgSO}_4 \cdot 7 \text{H}_2\text{O}$
Flowers of sulphur	S
Glauber's salt	$\text{Na}_2\text{SO}_4 \cdot 10 \text{H}_2\text{O}$
Glucose (dextrose)	$\text{C}_6\text{H}_{12}\text{O}_6$
Glycerine (glycerol)	$\text{C}_3\text{H}_5(\text{OH})_3$
Grain alcohol (ethyl)	$\text{C}_2\text{H}_5\text{OH}$
Graphite	C
Gypsum	$\text{CaSO}_4 \cdot 2 \text{H}_2\text{O}$
Household ammonia	NH_4OH
Hypo (crystals)	$\text{Na}_2\text{S}_2\text{O}_3 \cdot 5 \text{H}_2\text{O}$
Lampblack	C
Laughing gas	N_2O
Lime (hydrated)	$\text{Ca}(\text{OH})_2$
Lime (quicklime)	CaO
Lime (slaked)	$\text{Ca}(\text{OH})_2$
Limestone	CaCO_3
Limewater	$\text{Ca}(\text{OH})_2$
Lye	NaOH
Marble	CaCO_3
Marsh gas (methane)	CH_4
Methyl alcohol (methanol)	CH_3OH
Milk of magnesia	$\text{Mg}(\text{OH})_2$ (a suspension)
Moth balls (naphthalene)	C_{10}H_8
Muriatic acid (hydrochloric)	HCl
Nitre	KNO_3
Nitroglycerine	$\text{C}_3\text{H}_5(\text{NO}_3)_3$
Oil of vitriol	H_2SO_4

COMMON NAME	FORMULA
Plaster of Paris	$(\text{CaSO}_4)_2 \cdot \text{H}_2\text{O}$
Quicklime	CaO
Quicksilver (mercury)	Hg
Red lead	Pb_3O_4
Sal ammoniac (ammonium chloride)	NH_4Cl
Sal soda	$\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$
Saltpetre	KNO_3
Sand (silica)	SiO_2
Slaked lime	$\text{Ca}(\text{OH})_2$
Soap (sodium stearate)	$\text{C}_{17}\text{H}_{35}\text{COONa}$
Soda ash	Na_2CO_3
Sugar (sucrose)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}$
Superphosphate of lime	$\text{Ca}(\text{H}_2\text{PO}_4)_2 + 2 \text{CaSO}_4$
Vinegar (acetic acid)	$\text{HC}_2\text{H}_3\text{O}_2$
Washing soda (crystals)	$\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$
Wood alcohol	CH_3OH

M. FIRST AID IN THE CHEMISTRY LABORATORY

Faulty technique is one of the chief causes of accidents and, because it involves the human element, is one of the most difficult with which to cope. Good technique demands that students maintain an awareness of the danger in handling chemicals, are able to assemble apparatus correctly, handle materials properly, and proceed methodically. Proper methods of mixing acids, cutting glass tubing, and inserting it through stoppers, etc., are as important as correct experimental results. In spite of proper precautions, accidents do occur from time to time, and the fundamental rule to remember is to report the accident to the teacher immediately. Make certain that the right person begins first aid, and in the case of serious accidents, have *one* person call the school nurse or a physician immediately.

A supply of first-aid materials, kept in an easily accessible cabinet or cupboard, should be standard equipment in every laboratory. The following items are of value:

FIRST-AID MATERIALS

Rolls of 1", 2", and 3" sterile gauze bandage
 Sterile gauze, 3 inch squares.
 Absorbent cotton, sterilized
 Adhesive tape—to keep bandages in place

Band-aids
Eye cup
Scissors, tweezers, wood applicators, medicine droppers
Tincture of iodine (3·5% solution)
Petroleum jelly (vaseline)
Sodium bicarbonate solution (5%)
Boric acid solution (20 gm. in 1 litre of water)
Boric acid ointment
Butesin picrate ointment
Carron oil (equal volumes of limewater and linseed or cottonseed oil)
Castor oil or olive oil
Aromatic spirits of ammonia
Tincture of merthiolate
Naphtha gasoline or grain alcohol
Fire blanket

BURNS

Acid in the eyes. Pull back the lids, and flood the eye thoroughly with running water (a drinking fountain is excellent if immediately available). Then bathe the eye with a 5% sodium bicarbonate solution, using an eye cup. Dry with sterile gauze, and treat the eye with a few drops of olive oil or castor oil which acts as a soothing agent.

Base (alkali) in the eyes. Flood the eye thoroughly with running water. Then bathe the eye with a 5% boric acid ("boracic acid") solution, using an eye cup. Dry with sterile gauze, and add a few drops of olive oil or castor oil.

Acid burns on the skin. Wash the burned area thoroughly with running water, and then with a 5% solution of sodium bicarbonate. Apply sterile petroleum jelly or carron oil or butesin picrate ointment, and protect with gauze bandage.

Alkali burns on the skin. Wash thoroughly with running water, and neutralize with a boric acid solution. Wash again, and apply boric acid ointment or sterile petroleum jelly. Bandage the burned area loosely.

Heat burns. Cover the burned area with butesin picrate ointment or petroleum jelly, and then bandage loosely with sterile gauze. *Do not break any blister formed.* *Do not* use tincture of iodine on *any burn* because of its tendency to increase skin damage.

Bromine Burns. Wash thoroughly with running water, and immediately treat with a 5% solution of sodium bicarbonate. Dry, and apply carron oil or keep the burned area moist by applying glycerine.

Phosphorus Burns. Flood with water. Then apply a 1% aqueous solution of cupric sulphate. If no phosphorus remains in the burn, treat as for "heat burns" described above. If the burn is extensive, consult a physician.

INHALING TOXIC FUMES

Remove the victim to outside air, and encourage him to take deep breaths of fresh air. *Acid vapours* (chlorine, bromine, hydrogen chloride, etc.) may be treated by having the patient inhale the vapour from dilute ammonium hydroxide (aromatic spirits of ammonia). If the gas is *ammonia*, allow him to inhale the vapour from a dilute solution of acetic acid.

Headaches or dizziness are nature's warning of toxic atmospheres. Fresh air, five grains of aspirin, and a half hour of rest will usually relieve toxic headaches.

CUTS

If bleeding is copious, it must first be stopped. An arterial cut is indicated by scarlet blood flowing intermittently. Use a piece of rubber tubing (half inch bore) as a tourniquet between the heart and the cut. Never tighten more than necessary to stop the bleeding, and do not maintain the pressure for longer than fifteen minutes at a time. Loosen it and allow blood to flow for a few seconds, and then retighten. Loosen, but do not remove the tourniquet as the blood clots. **Send the student to a physician.**

The blood which pours from veins is dark red in colour and appears in a steady flow from the wound, without the pulsations or "spurts" seen in an arterial cut. Venous bleeding is more readily controlled. Apply a sterile gauze compress directly over the wound, and apply firm pressure with the fingers until the bleeding stops or a clot forms. When the bleeding ceases, leave the compress undisturbed and bandage with sterile gauze.

Minor cuts are serious because of possible ingress of bacteria, leading to infection. Wash the wound zone with a gauze compress saturated with naphtha gasoline or grain alcohol to remove any grease. If foreign particles are present on the surface, they should be removed by gentle brushing with sterile gauze. Apply an antiseptic solution (tincture of iodine, tincture of merthiolate, mercurochrome solution, etc.), and bandage with sterile gauze.

Patients with deep puncture wounds, as from glass tubing, broken thermometers, etc., **should be sent to a physician.**

ACCIDENTS TO CLOTHING

Acids on clothing. Apply limewater or a dilute solution of sodium bicarbonate to the affected area to neutralize the acid. Then wash well in running water.

Bases on clothing. Neutralize the alkali by applying dilute acetic acid to the affected area. Then wash thoroughly with water.

Clothing on fire. Smother the fire (exclude oxygen) with a woollen blanket. A suit coat may be used if the laboratory is not equipped with a fire blanket. Speed is essential if severe injury is to be avoided.

FAINTING

Place the patient in a comfortable lying position with the head lower than the rest of the body. Open the windows, and insist that onlookers do not crowd around the patient, since a good circulation of air is imperative. The return to consciousness can be accelerated by having the patient breathe the vapour from aromatic spirits of ammonia placed on a folded handkerchief. Placing a towel moistened with cold water on the patient's forehead is also helpful. When consciousness is restored, the patient should be kept lying down until fully recovered, and should then be allowed to leave the classroom only when accompanied by reliable assistants.

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THE ELEMENTS

ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT	ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT
Actinium	Ac	89	227	Mercury	Hg	80	200.61
Aluminum	Al	13	26.98	Molybdenum	Mo	42	95.95
Americium	Am	95	(243)	Neodymium*	Nd	60	144.27
Antimony	Sb	51	121.76	Neon	Ne	10	20.183
Argon	A	18	39.944	Neptunium	Np	93	(237)
Arsenic	As	33	74.91	Nickel	Ni	28	58.71
Astatine	At	85	(211)	Niobium	Nb	41	92.91
Barium	Ba	56	137.36	Nitrogen	N	7	14.008
Berkelium	Bk	97	(245)	Nobelium	No	102	(253)
Beryllium	Be	4	9.013	Osmium	Os	76	190.2
Bismuth	Bi	83	209.00	Oxygen	O	8	16.0000
Boron	B	5	10.82	Palladium	Pd	46	106.7
Bromine	Br	35	79.916	Phosphorus	P	15	30.975
Cadmium	Cd	48	112.41	Platinum	Pt	78	195.09
Calcium	Ca	20	40.08	Plutonium	Pu	94	(242)
Californium	Cf	98	(248)	Polonium	Po	84	210
Carbon	C	6	12.011	Potassium	K	19	39.100
Cerium*	Ce	58	140.13	Praseodymium*Pr	Pr	59	140.92
Cesium	Cs	55	132.91	Promethium	Pm	61	(145)
Chlorine	Cl	17	35.457	Protactinium	Pa	91	231
Chromium	Cr	24	52.01	Radium	Ra	88	226.05
Cobalt	Co	27	58.94	Radon	Rn	86	222
Copper	Cu	29	63.54	Rhenium	Re	75	186.22
Curium	Cm	96	(245)	Rhodium	Rh	45	102.91
Dysprosium*	Dy	66	162.51	Rubidium	Rb	37	85.48
Einsteinium	E	99	(255)	Ruthenium	Ru	44	101.1
Erbium*	Er	68	167.27	Samarium*	Sm	62	150.35
Europium*	Eu	63	152.0	Scandium	Sc	21	44.96
Fermium	Fm	100	(252)	Selenium	Se	34	78.96
Fluorine	F	9	19.00	Silicon	Si	14	28.09
Francium	Fr	87	(223)	Silver	Ag	47	107.880
Gadolinium*	Gd	64	157.26	Sodium	Na	11	22.991
Gallium	Ga	31	69.72	Strontium	Sr	38	87.63
Germanium	Ge	32	72.60	Sulphur	S	16	32.066
Gold	Au	79	197.0	Tantalum	Ta	73	180.95
Hafnium	Hf	72	178.58	Technetium	Tc	43	(99)
Helium	He	2	4.003	Tellurium	Te	52	127.61
Holmium*	Ho	67	164.94	Terbium*	Tb	65	158.93
Hydrogen	H	1	1.0080	Thallium	Tl	81	204.39
Indium	In	49	114.82	Thorium	Th	90	232.05
Iodine	I	53	126.91	Thulium*	Tm	69	168.94
Iridium	Ir	77	192.2	Tin	Sn	50	118.70
Iron	Fe	26	55.85	Titanium	Ti	22	47.90
Krypton	Kr	36	83.8	Tungsten	W	74	183.86
Lanthanum*	La	57	138.92	Uranium	U	92	238.07
Lead	Pb	82	207.21	Vanadium	V	23	50.95
Lithium	Li	3	6.940	Xenon	Xe	54	131.30
Lutecium*	Lu	71	174.99	Ytterbium*	Yb	70	173.04
Magnesium	Mg	12	24.32	Yttrium	Y	39	88.92
Manganese	Mn	25	54.94	Zinc	Zn	30	65.38
Mendelevium	Mv	101	(256)	Zirconium	Zr	40	91.22

*Rare Earths

Atomic weight in brackets denotes the mass number of the isotope of longest known half-life.

APPROXIMATE ATOMIC WEIGHTS OF THE COMMON ELEMENTS

NAME	SYMBOL	ATOMIC WEIGHT	NAME	SYMBOL	ATOMIC WEIGHT
Aluminum	Al	27	Lead	Pb	207
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Carbon	C	12	Phosphorus	P	31
Chlorine	Cl	35.5	Potassium	K	39
Copper	Cu	63.5	Silver	Ag	108
Fluorine	F	19	Sodium	Na	23
Gold	Au	197	Sulphur	S	32
Hydrogen	H	1	Tin	Sn	119
Iodine	I	127	Zinc	Zn	65
Iron	Fe	56			

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